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The World of Matter

A GUIDE TO THE STUDY OF

CHEMISTRY AND MINERALOGY

BY

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President of the Agassiz Association, Author of "Three Kingdoms," "American Plant Book," "One Thousand Blunders in English Corrected," etc.

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THIS BOOK

IS

AFFECTIONATELY DEDICATED

TO

MY MOTHER.

PREFACE.

In 1875 the author founded a society for the study of Natural Science at home, which, under the name AGASSIZ ASSOCIATION, has during sixteen years enlisted no less than twenty thousand students of all ages, and organized more than one thousand local branch societies scattered throughout the world. Under the auspices of the Agassiz Association and under the general supervision of the author, who has heretofore conducted the entire management and correspondence of the society, annual courses of experimental study have been conducted by eminent teachers, through the mails. During the present year nearly one thousand pupils in the United States are at work on a course in Mineralogy under the care of Professor Gustave Guttenberg.

This practical experience of sixteen years has firmly convinced the author that the natural sciences are ordinarily taught in a fragmentary manner. Wonderful advances have indeed been made in methods of teaching them within twenty years, and there is now a multitude of competent and faithful teachers, who, each in his own special department, are training their pupils as

they ought to be trained, discarding ancient systems of memorizing and rote-work, and introducing laboratory and field-work, which develops and strengthens the natural powers of perception and reason: but, rarely, if ever, has the attempt seriously been made so to co-ordinate the several branches of science as to cause the student to apprehend their essential unity, or even to approach them in a natural and logical order.

To accomplish this, is attempted in this series of books. A brief discussion of the true order of scientific studies is included in Chapter XIX., and need not be repeated here.

No one need hope to learn anything of much value either of chemistry or mineralogy unless he handles specimens and makes experiments.

A collection of thirty-six minerals has been prepared to accompany this book. They are arranged in a neat case. They have been selected under the supervision of Prof. W. O. Crosby, of the Boston Society of Natural History, and can be depended upon as excellent typical specimens, and they are furnished at the bare cost of the labor required in their preparation. Nevertheless it is not necessary that the reader should procure these particular specimens, as he will be able to find similar ones almost anywhere.

The author has made free use of the standard text-books in Physics, Chemistry, and Mineralogy, particularly Ganot, Roscoe and Schorlämmer, Cooke, Faraday, Huxley, Tenney, Richards, Kunz, and Dana; he is especially indebted to the generous courtesy of the

authors and Publishers of Remsen's Elements of Chemistry, Crosby's Common Minerals and Rocks, and Guttenberg's Course in Mineralogy; and he desires to express his gratitude to Professors Ira Remsen, W. O. Crosby, Gustave Guttenberg, and John B. Welch, for valuable suggestions during the progress of the work, and for reading the proofs while in press.

PITTSFIELD, MASS., September 15th, 1891.

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CONTENTS.

Chapter.		Page.
	Preface	iii.
I.	A Piece of Ice	7
II.	A Piece of Ice (Continued)	16
III.	A Glass of Melted Ice	33
IV.	Is Water an Element?	45
V.	Fire	55
VI.	Air	67
VII.	Earth	77
VIII.	Quartz	86
IX.	A Lesson in Chemistry	94
X.	A Piece of Charcoal	103
XI.	Carbon Continued	115
XII.	A Piece of Marble	126
XIII.	Clay	137

CONTENTS.

Chapter.		Page.
XIV.	Potassium—Mica	144
XV.	A Lump of Salt	150
XVI.	Muriatic Acid	156
XVII.	Chlorine	161
XVIII	Iron	165
XIX.	By Way of Review	171
XX.	What is a Metal?	179
XXI.	Families of Acid-Forming Elements	184
XXII.	Families of Metals	194
XXIII	How to Determine Minerals	198
XXIV.	Flame Tests, Heat Tests and Acid Tests	205
XXV.	Twelve Other Minerals	216
XXVI.	Blow-Pipe Analysis	226
XXVII.	Suggestions for Further Study	245

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THE WORLD OF MATTER.

CHAPTER I.

A PIECE OF ICE.

Ice is one of the most common minerals. Let us begin our study with a piece of ice. We will not ask any one to tell us anything about it until we have learned what we can from our own observation.

Common sense is the sense which we all have in common and by the use of our common sense, and our common senses, we can learn enough for practical purposes. The senses which we all have are feeling, sight, taste, hearing, and smell. What do these teach us about a piece of ice? We notice first that it is cold, smooth, slippery, hard, and heavy. We learn all this from the sense of feeling by holding the piece of ice in the hand.

No scientist understands the meaning of these words better than we, and no explanation can make them any plainer. No facts in natural science are any more difficult to understand than these. The only reasons for the popular notion that there is something hard and mysterious about chemistry and mineralogy are that many of the substances treated of in the books are less familiar to us than ice, and that many of the words used to describe them are new to us. When a piece of muscovite

becomes as familiar to our hand as a piece of ice, and when we become as familiar with the words "transparent" and "foliaceous" as we are with the words "hard" and "smooth," the statement that "Muscovite is transparent and foliaceous" will seem as simple as that "Ice is hard and smooth."

This is by the way, but it is important, for I wish the reader to remember that if he ever finds anything in a good scientific text-book which seems hard to understand, it will be because he is not familiar with the thing described, or because he does not know the meaning of the words used to describe it, or because he has not mastered the preceding portion of the book.

To remedy the first trouble it is necessary to get hold of the strange substance and examine it; to remedy the second trouble the strange words must be clearly defined.

Let us return to our piece of ice, and having learned what our hands can tell us of it, let us see what our eyes can teach us. This piece is thin and clear. It looks like a piece of glass. If we lay it on this page we can read the letters through it. If it were large and strong enough it would do for a window-pane. The light comes through it. If we hold it in the sunlight it shines like glass.

Testing it now by our other senses we find that it has no smell or taste, and that when we break it, it breaks with a snapping or cracking sound.

In this connection we notice that it is brittle, a fact to which our hand, eye and ear all bear witness.

We have spoken of our piece of ice as a "thing," but it is better to use the word "substance." This word is better because the word "thing" is too indefinite, being used also of events or deeds.

Shall we say then that ice is a brittle, hard, smooth, slippery, clear, shining, heavy, and cold substance, without taste or smell? Not yet. All this is true of the particular piece we are examining, but we remember that we have seen other pieces of ice that were rough instead of smooth, dull rather than shining and white or green instead of clear. We must therefore recollect that we are speaking only of a particular sample of ice, not of ice in general. Such a sample of any substance is called a "specimen;" and it is only after a study of many specimens, and after finding what is true of all of them, that we shall be ready to make a general description or definition of ice.

In all the ice we have seen we have observed coldness, brittleness, hardness, and weight. Coldness belongs to ice wherever it is found, and therefore is a "property" of ice, as whatever belongs to a man is his property. Brittleness, hardness, and weight are also properties of ice, and as these have all been perceived through the senses, they are called "physical" properties.

The first step in the study of any substance is the determination of its physical properties.

A set of mineral specimens has been prepared to accompany this book, and in the case of each of them you will do well to make a list of its physical properties as revealed by your senses, as we have with our

piece of ice. The names for such additional properties as may be observed will be given and explained as we go on.

If you have held your specimen of ice in your hand while examining it, you must by this time have observed a remarkable change in it. It has partly melted. Its edges are no longer sharp but rounded; it feels wet, and water is trickling from it. In a short time there will be only a few spoonfuls of water instead of a piece of ice. The specimen has changed from a solid to a liquid condition. This change has been a silent one and has been learned through the two senses, feeling and sight.

Up to this point our study has been entirely by observation. We have simply examined our specimen and observed such of its physical properties as have revealed themselves to our senses without any special effort on our part. It is now time to go a step further and make an experiment. This consists in placing our specimen under new conditions and observing how it behaves. If we had placed the ice in the hand with the intention of learning whether any effect would be produced thereby, and if we had then observed that it melted, that would have been an experiment. Experiments are made in order to discover something that we cannot learn while our specimen remains under ordinary conditions; and they are very frequently made in order to discover the cause of some appearance or change which we have noticed. Any such appearance or change is called in scientific language a "phenomenon."

Whatever is perceived by any or all the senses is a phenomenon. Among the more common phenomena are light and darkness, heat and cold, motion and rest.

The melting of ice is a phenomenon, and having observed it, let us make a few simple experiments in order that we may if possible discover its cause. The fact that in this case we already know the cause, is no hindrance to the experiments. They will only confirm us in our present opinion if that is right. First then lay a piece of ice, wrapped in flannel, in a very cold refrigerator. This is the experiment. The observation is that it does not melt. Next place it near a hot fire. It melts. Try it in hot and cold water. In the warm rays of the sun. Hold it in your hand or your mouth. If then, there is any one condition, and only one, which is present in every instance when the ice melts, and always absent when it does not melt, it will be reasonable to infer that whatever produces that condition is the cause of the phenomenon of melting. Now, in each experiment the condition of melting is that something warm be near the ice; either warm air, or a warm hand, or warm iron, or warm water; and as nothing is made warm without heat, we arrive at the conclusion that heat causes ice to melt. This property of ice, namely, that it melts when heated, is termed "fusibility." Ice is fusible.

Our minds are so constituted by nature that we invariably look for some cause for every observed phenomenon, and it is a curious fact that even before we learn with certainty the true cause, we can hardly

refrain from risking a guess at it as soon as the phenomenon is observed. It is nearly necessary to make such a guess before we can undertake an intelligent experiment, and experiments are usually made with the purpose of determining the correctness or incorrectness of a previous guess at the true cause. Such a guess, especially when made the basis of experiments, is called an "hypothesis," and hypotheses are of the highest value in all scientific study.

An hypothesis is the first answer the mind suggests to the question "Why?"

It must not be confounded with a theory, which is a full and orderly explanation of a group of observed facts and phenomena, and usually accounts for a large number of facts apparently different and unconnected. A theory often rests on an hypothesis; as the Darwinian theory of the origin of species depends largely upon the hypothesis that accidental variations may result in forms sufficiently new and permanent to be reckoned as distinct species.

Having learned that heat causes ice to melt, the curious student will wish to know what causes the heat; and having if possible discovered that, he will again search for the cause of that cause, and so on in an ever lengthening chain, until he either reaches some first cause, or becomes weary or discouraged in its pursuit. To this first cause religion makes a short path by asserting the existence of an infinite and eternal being, God. Toward it science moves more slowly, following back the chain of effect and cause link by link, and verifying

each conclusion by patient and prolonged experiment. Some scientists, like Louis Agassiz, ultimately arrive at the same theological conclusion—a creator, God; others find a first cause in a universal and unintelligent force constantly acting under natural law; others declare that the first cause of all natural phenomena is so remote as to be wholly beyond the reach of human knowledge or imagination; and others are contented to work as faithfully as may be along the nearer and clearer links in the great chain, without speculating upon that part of it which is out of sight. For the present we shall remain among this latter class.

The question of the nature and cause of heat will be reserved for another chapter.

Up to this point our observations and experiments have been of the simplest possible character, and have consisted merely in observing those marked physical properties of ice which at once present themselves to the most careless mind. Let us now make a somewhat more careful examination. Taking another specimen as nearly like the one we have melted as possible, we not only perceive that it is hard, but will now raise the additional question, "how hard?"

In order to test the degree of hardness which any mineral possesses, it becomes necessary to have a measure or scale of hardness. That is, as in the case of all measuring, we must have a known and familiar standard with which to compare each new specimen. The scale adopted by mineralogists consists of a series of ten specimens, beginning with one of the softest, talc,

whose hardness is called 1, and ending with the hardest known mineral, diamond. The scale is as follows:

SCALE OF HARDNESS.

1. Talc.	6. Orthoclase.
2. Gypsum.	7. Quartz.
3. Calcite.	8. Topaz or beryl.
4. Fluorite.	9. Corundum.
5. Apatite.	10. Diamond.

Specimens of the first seven of these are included in the collection which accompanies this book. They are easily obtained, and as minerals harder than quartz are very few and very rare, they are all that you really need. We test the relative hardness of two minerals by finding out which will scratch the other. If any mineral, for example, will make a distinct scratch upon a piece of talc, it is harder than talc; unless, indeed, it can itself be scratched as distinctly by a piece of talc, in which case they may be considered to have the same degree of hardness. The reason that minerals of the same degree of hardness can scratch each other is that a sharp corner has a slight advantage over a smooth surface. You can often scratch a pane of glass with a fragment broken from one corner of it.

There is a simpler method of testing the hardness of minerals, and a method sufficiently accurate after a little practice. The scale of hardness in this case consists of your thumb-nail, a pocket knife, and a piece of ordinary window-glass. Specimens that can be scratched with the nail as easily as talc, have a hardness expressed by the

number 1; those that offer more resistance, but yet can be scratched, like gypsum, have a hardness of 2; those that can very easily be scratched with your knife-blade have the 3d degree of hardness; those that require more pressure under the blade, have hardness 4; those that can be scratched with a knife with difficulty, and are not yet hard enough to make a scratch on glass, have hardness 5; those that make a slight scratch on glass, and are yet soft enough to be scratched by the knife-blade with the greatest difficulty, have hardness 6; while those that easily scratch glass and resist the edge of the blade, have hardness 7 or more.

It must not be considered a proof of the softness of a mineral that it crumbles under pressure. Sandstone which you can crush between your fingers is often composed of particles of quartz which will scratch glass. Test the hardness of your specimen of ice.

QUESTIONS ON CHAPTER I.

1. What is ice?
2. Name four of its physical properties.
3. What is a specimen?
4. What change takes place in a piece of ice held in the hand?
5. What is the cause of this?
6. What is the difference between observation and experiment?
7. What is a phenomenon?
8. What is an hypothesis?
9. What is a theory?
10. Explain the "scale of hardness."

CHAPTER II.

A PIECE OF ICE—(CONTINUED.)

At our first examination of a piece of ice we observed that it was heavy. Let us now determine how its weight compares with the weight of other mineral specimens of the same bulk. This is one of the most important points in the study of any specimen. Here, again, as in the case of hardness, we must have a standard of comparison; that is, some substance with whose weight we are familiar, with which we may compare each new specimen we meet. From its abundance and universal distribution, water is taken as the standard. The weight of water varies, however, with its temperature, pressure and purity, and when extreme accuracy is required, we use distilled water, weighed at a fixed temperature—commonly 60 degrees Fahrenheit—and under such atmospheric pressure as raises the mercury in a barometer to the height of 30 inches: but for our present purpose we may neglect the questions of temperature and pressure, and even purity, and simply compare the weight or density of our specimen with that of an equal bulk of whatever water is most convenient to our hand. The relative weight of any substance as compared to that of water

is called its specific weight or specific gravity. If I say that the specific gravity of cast zinc is 7, I mean that a piece of the zinc is seven times as heavy as an equal bulk of water.

Students commonly find something difficult or confusing about this term "specific gravity," because the words are unusual. If you do not perfectly understand it, take one or two more illustrations. If a cubic inch of clay is twice as heavy as a cubic inch of water, its specific gravity is 2. The specific gravity of melting ice is .93, and this means that a cubic inch or a cubic foot of ice weighs $\frac{93}{100}$ as much as a cubic inch or foot of water.

The word "specific" means "belonging to a particular kind;" so that if we translate the expression "the specific gravity of a mineral" into common words, it reads "the weight of a particular kind of mineral;" but we must remember that this does not mean its absolute weight as determined by the scales, but its relative weight as compared to an equal bulk of water.

The question now arises how the specific gravity of any specimen can be determined. If our specimen happened to be of such a shape and size as exactly to fit into any convenient vessel, as a cup, it would be a very simple matter. We should then fill the cup with water, weigh it, and subtract the weight of the empty cup. This would give the weight of the water. Then we should weigh our specimen; and if we found, for example, that the water weighed 8 ounces, and the specimen 30, we should know that the specimen was

$3\frac{3}{4}$ times as heavy as an equal bulk of water, or that its specific gravity was 3.75.

What shall we do when, as is nearly always the case, our specimen is irregular in shape?

We might make a mould in plaster-of-paris, which would hold the same bulk of water as the specimen, and then weigh and compare as before; but this would be tedious, and inexact, particularly if our specimen happened to be porous or hollow.

We might fill a vessel with water quite up to the level of a waste-pipe, as in Fig. 1.

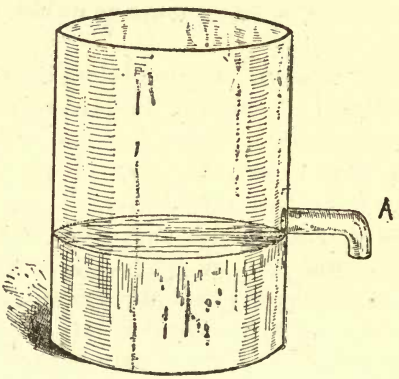

Fig. 1.

If, then, we lowered the specimen into the water it would displace an equal bulk of water which, flowing out at A, could be caught and weighed.

This also would be a slow and awkward process, and liable to error.

Just at this point we are greatly helped by a fact

which was discovered about 1600 years ago by a famous philosopher of Sicily, named Archimedes. The interesting and amusing story of his discovery you will find in any cyclopedia, but the fact is this, that when anything is plunged into water it seems to lose just as much weight as the weight of the water it displaces. Now, of course, any specimen plunged into water displaces an amount of water equal to its own bulk. Therefore if we ascertain how much loss of weight a specimen sustains in water, we know exactly how much an equal bulk of water weighs. For example, lower a piece of zinc weighing in the air seven pounds into water and weigh it again. It now weighs only six pounds. This means that it has displaced exactly one pound of water; or that water equal in bulk to 7 pounds of zinc, weighs one pound; or that zinc is seven times as heavy as water; or that the specific gravity of zinc is 7.

To find the specific gravity of any mineral, therefore, *divide its weight in air by its loss of weight in water.* Strictly speaking, we should use the weight of the specimen in a vacuum instead of in air, but that is not necessary for practical purposes. Convenient balances for this air and water weighing are furnished by all dealers in chemical apparatus, and need not be described here, particularly as you will not need to use them until you have advanced further in your work. For the present, after a little practice, and a careful handling of equal-sized pieces of the following substances, which constitute a sort of specific gravity

scale of approximate exactness, you can estimate the specific gravity of a mineral pretty closely by lifting it in the hand.

SPECIFIC GRAVITY SCALE.

Seasoned live oak, or ice, 1.
Gypsum, or sulphur, 2. (Common feldspar, 2.5.)
Cryolite, or apatite, 3.
Corundum, 4.
Magnetite, or pyrite, 5.
Cuprite, 6.
Cast zinc, 7.
Bell metal, 8.
Cinnabar, when pure, 9. (Copper nearly 9.)
Silver, 10.5.
Gold, or platinum, 20.

We have already observed that ice shines in the light. This property is termed lustre, and the various kinds of lustre are named from substances in which these varieties are most commonly observed; namely, glass, pearl, metal, resin or wax, diamond, and silk. Lustres resembling these are called vitreous or glassy, pearly, metallic, resinous or waxy, adamantine, and silky. The two most important are vitreous and metallic. Indeed a chief division of all minerals is into two classes;

1. Those that have a metallic lustre. 2. Those that have not. If a mineral has no lustre at all it is called dull.

We observed that our specimen of ice was clear, that we could see to read through it. This property is called

"transparency." Air, water and glass are transparent. Minerals through which one can perceive light, like ground glass or smoked glass, but through which one cannot distinguish objects clearly, are said to be translucent; those through which no light at all can be seen, like the metals, are called opaque.

We observed that our specimen was nearly colorless, yet we recollected that we had seen ice that, instead of being clear, was white and green. This leads us to notice the distinction between the real, or essential color of a mineral, and its apparent color. The apparent color of minerals is one of their least important properties, because it arises from so many different causes that it is very variable. This variability is more marked in transparent than in opaque substances, and often arises from some impurity. Thus quartz may be white, yellow, red, green, blue, brown, or black, the color in each instance being due to the presence of some other mineral diffused throughout the quartz. A very minute quantity of coloring-matter is sufficient to change the hue of a large mass, just as a few grains of indigo make a large tank of water blue.

The color of the powder of any mineral is much less likely to be affected by the presence of these impurities, and we therefore regard that as of far more importance than the apparent color in identifying it. This powder may be obtained by crushing a small fragment, by scratching the specimen, or by making a mark with the specimen upon a surface harder than itself. An ordinary slate-pencil is apparently of a dull lead color-

but crush a fragment to powder, or scrape it with a knife, or observe the color of the mark it leaves upon the slate, and you will learn that its essential color is light gray, approaching white.

The simplest way to observe this essential color, in the case of most minerals, as in the case of the slate pencil, is by observing the color of the mark or "streak" it makes upon a harder surface. For this reason the plate of hard glass generally used to test this color, is called a "streak-plate," and the color itself is called the "streak." Now whether the apparent color of ice be green or blue or white, the color of its powder or "streak" is always white. This is the case also with nearly all transparent and translucent minerals, whatever their apparent color may be. The real color of ice appears most conspicuously in that form of it known as snow.

We have already observed that ice is easily fusible, and we have learned by trial that the cause of its melting is heat. Let us now try a further experiment with a view to determining the degree of heat necessary. To test this approximately we have only to place the bulb of a thermometer in a mixture of melting snow or ice, and we shall find that it registers 32 degrees Fahrenheit. We reach the same result by placing the thermometer in water which is gradually cooling, and reading it at the moment when freezing begins, for the freezing-point and the melting-point are so near together as to be practically the same. It is earnestly hoped that the student will not be satisfied with reading about these experi-

ments, but that he will actually make them, no matter how trivial they may appear. Otherwise most of the advantage of this book will be lost. Little facts are constantly revealing themselves to the eye of the experimenter of which no mention is made in the most elaborate book. Books give general statements. They record the results of experiments made by the most skillful men, with the most delicate apparatus, and under the most accurately defined conditions.

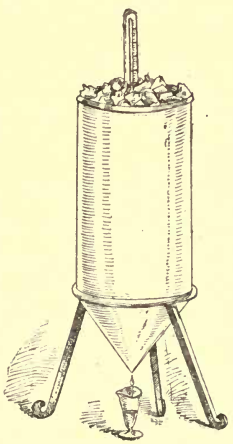

Fig. 2.

Do not, therefore, accept without experiment our statement that ice melts or water freezes at a temperature of 32 degrees. Try it with your own thermometer. The chances are that no ice will appear on your cup of water until the mercury has dropped to 30 degrees, perhaps even to 28 degrees Fahrenheit. Should this be the case it would result either from the presence of some impurity, as salt, in the water, increase of pressure or on account of its perfect stillness. In the latter case a slight agitation of the water will cause part of it to shoot suddenly into crystals of ice, while the temperature of the remainder will rise to 32 degrees.

A convenient apparatus for testing the freezing-point of water is shown in Figure 2.

Snow or pounded ice is placed in this vessel, from

which water escapes below. The bulb and part of the tube of a thermometer are immersed in this for about a quarter of an hour, and then the mercury should stand at 32°; for although the freezing-point of pure water can be retarded, as has been shown, its melting point is always the same, or at least much less variable, as it requires great increase of pressure to lower it perceptibly. In this connection we must notice two very important phenomena which attend the cooling and freezing of water, namely contraction and crystallization.

EXPERIMENT. Provide a glass vessel having a slender tube rising from a large bulb, as in Figure 3. Insert a thermometer in this vessel. The scale can be read through the glass.

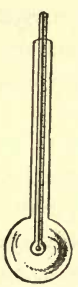

Fig. 3.

Fill the bulb, and half the tube with water and gradually heat it. If the water rises in the tube it will show that it is expanding. The larger the bulb in proportion to the stem the more perceptible, and apparently rapid, will be the effect. Now surround the bulb with a mixture of pounded ice and salt and note the effect of cooling the water. By all means make this experiment yourself, as it forcibly illustrates the danger of hastily coming to conclutions that may be erroneous. Thus if you perform this experiment carelessly, you will almost certainly decide that water always expands when heated and contracts when chilled. It does, so far as you have observed, and you see no reason why it should ever act otherwise.

But as the water in cooling approaches the freezing-point you must watch it more attentively. The unexpected often occurs in nature. We constantly run upon real or apparent exceptions to what we have imagined to be an invariable rule. The water in the tube keeps on contracting just as you would expect, until it becomes cooled to a temperature of about 39° Fahrenheit, but at that point the contraction suddenly ceases, and expansion begins. The column of cooling water in the tube stops falling and begins to rise again. This degree of temperature, therefore, 39°, is the point of the greatest relative weight of water; or, to speak scientifically, water at 39° has its maximum density. If you find any difficulty in observing this phenomenon in the simple tube just now described, you can repeat an experiment first made by Thomas Charles Hope, a professor of chemistry in the universities of Edinburgh and Glasgow many years ago.

"Insert two thermometers (Fig. 4,) at different levels into a cylinder of water, and chill the water by applying ice around the middle of the vessel. As the water becomes cooled it grows denser, and therefore sinks to the bottom, so that the *lower* thermometer falls until it reaches 39° Fahr. Further cooling then *expands* the water, instead of condensing it, and consequently the cold water rises, so that now the upper thermometer, which has meanwhile been almost stationary, begins to fall, and continues falling, until, like the lower one, it reaches 39° Fahr. The whole body of water is then at its maximum density, and any further reduction of

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temperature causes expansion, the cold water becoming specifically lighter and rising to the surface. Gradually the upper thermometer sinks to the freezing point, and then a layer of ice begins to form upon the surface. This experiment roughly imitates what occurs in a natural piece of water, such as a lake; the surface freezes, while the bottom water remains several degrees warmer."

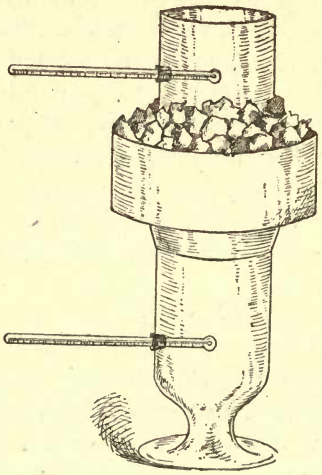

Fig. 4.

This experiment, while not quite so simple as the one first described, has the advantage that it makes clearer the reason of the very important fact that water naturally freezes over first at the top. Were it not for this strange property of water, by virtue of which it stops contracting and growing heavier just be-

fore it reaches the freezing point, northern lakes and rivers would freeze in winter into solid masses of ice, which would destroy all aquatic life. Moreover, the heat of summer would scarcely suffice to melt them, and the climate might in time be changed into a perpetual winter such as now lingers about the poles.

The increase in the bulk of water, which begins at 39° Fahr., continues until the moment of freezing, when a sudden and much greater expansion occurs with almost irresistible force.

Fill the strongest vessel you can get, even a cannon or hollow bombshell, with water; plug it tight, and let the water freeze: you will have ample evidence of this force.

Do not now jump to the conclusion that all liquids behave in the same way as water under like conditions, and do not consult books in order to find out, but if you care to know, test them carefully yourself, one by one.

One experiment devised and performed by yourself may be of more educational value than a year's study of text-books.

The ice which forms upon water appears at first sight to be a solid mass without definite structure. It is really, however, built up of innumerable crystals, each of a definite shape, which are so interlocked and compacted together that their separate forms are obscured. Professor Tyndall succeeded some years ago in revealing this hidden architecture, by sending a beam of sunlight through a block of ice. Enough heat penetrates the ice from such a sunbeam to cause a slow melting inside the

block. This first manifests itself by the appearance of little glistening points, and then rays of light shoot out from these bright centres until the ice seems full of shining snowflakes. These beautiful forms have been called "ice-flowers." They are not really ice-crystals, however, but only the moulds left by the melted crystals, and now nearly filled with water, which reflects the light and discloses the secret of the structure of the ice. The glistening point in the centre of each liquid star is a vacuum or empty space left by the water which shrinks as it melts; and the surface of water as it curves about this vacuum shines in the sun like silver.

The beauty and symmetry of ice-crystals, which the solid ice kept for ages locked up in its frozen breast until Professor Tyndall opened it with his key of light, is nevertheless plainly shown to us in every shower of snow; when, if the air is still, each flake that falls is a perfect and exquisite crystal. You will be well repaid for your pains if you catch a number of these snow-crystals next winter upon a cold pane of glass, or a black cloth, and study them with a microscope. The variety of forms which they assume is infinite, and their beauty indescribable; yet it will be found that they are all fashioned on a definite plan, and have their delicate rays always arranged in threes or sixes. These six-angled or hexagonal flakes, however, are not regarded as single ice-crystals, but rather as many little crystals symmetrically united. I have never been able to see satisfactorily and clearly one of the tiny crystals of which these snow-stars are said to be composed, but the "books"

say that their form is that shown in Fig. 6; that is to say, a solid bounded by six equal lozenge-shaped surfaces. These diamond-shaped surfaces are called "rhombs," by the geometricians, and a solid of the form shown in Fig. 6 is called a "rhombohedron." Just how these tiny rhombohedrons are joined together or modified to produce the six-angled or "hexagonal" forms seen in snow-flakes and "ice-flowers" can not be clearly understood until you get glass or wooden models and experiment with them.

Fig. 6.

If you have a model of a rhombohedron, you will see that you can hold it so that an end view presents a six-sided outline; just such a form, in fact, as appears in the centre of a snow-flake. Now as these snow-flakes are usually transparent, always very thin and delicate, and as it is difficult to examine them with a glass of high magnifying power, it is not impossible that, instead of being flat, hexagonal prisms, as they are commonly supposed to be, they are really rhombohedrons, whose upper edges escape our observation; and it may afford you some interest to examine them for yourself next winter.

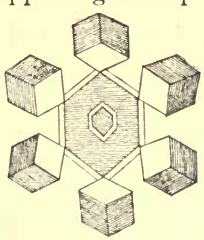

Fig. 7.

In one of the snow-crystals drawn at Greenwich, England, by Mr. Glaisher, and printed in Tyndall's "Forms of Water," the rhombohedral form of ice is clearly shown in the six crystals which surround the hexagonal centre, Fig. 7, and if we may suppose

Mr. Glaisher to have overlooked the delicate angles of the central portion of this snow-flake, which would only be revealed by the closest observation, we may supply them, as in Fig. 8; and then we have a compound crystal, symmetrically composed of regular rhombohedrons, instead of a combination of rhombohedrons with a central hexagonal prism.

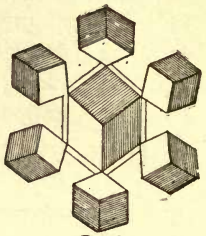

Fig 8.

When a very thin film of water freezes on a smooth surface, as the moisture on a window-pane, its crystals do not usually present to the eye either the hexagonal or rhombohedral form assumed by vapor that freezes in the air and falls as snow. They curve in marvellous imitations of trees and flowers; they build themselves into fairy castles with crystal towers and battlements; they shoot into twinkling spears and glistening needles; they curl into waving plumes of exquisite grace and beauty. It is fascinating to make photographs of this frost-work. It is not impossible that by a comparison and study of a large number of such photographs we might learn the secret of the curves and angles that seem so capricious in their loveliness. Do the shooting crystals follow the curve of some eddying current of air? Do they run along some slight irregularity in the structure of the glass? Are they modified by the nearness of nails or other forms of metal? Are they controlled at all by musical tones? Can they be affected by a gentle current of electricity? Here is an enticing field for observation and experiment.

A kind of ice-crystal less commonly noted occurs in moist ground, in the shape of long sharp blades and needles projecting an inch or more from the surface of the ground, and standing like elfin troops on drill, all armed with gleaming bayonets. Along the margins of brooks in winter are found still more wonderful forms of crystal architecture. Drooping ferns and curling leaves and waving blades of grass are reproduced along the edge of the shivering water, giving one the impression that the plants that fringed the stream in summer are reappearing in shadowy shapes of ice.

One December evening there came a chilling fog. Later the mercury fell rapidly, and a rush of cold followed, as if a window had been opened into the Arctic zone. In the morning the fog was gone, the sun shone bright, and the sky was blue; but every strip of exposed metal, and every edge of unpainted wood, was bordered by a fringe of soft white feathers ranged as closely together as the down on the breast of an eider-duck. On our front porch is a wire netting for summer vines to run on, consisting of hexagonal meshes two inches in diameter. Never clematis or ivy decked this slender trellis so gracefully as the vine of water-crystals on that icy morning. The sudden and unwonted beauty brought exclamations of delight to the lips of all who saw it; but before a camera could be found to preserve the picture, the sun silently undid the fastenings of the tiny plumes, and they dropped to the earth in showers of silvery sheen.

QUESTIONS ON CHAPTER 2.

1. What is the specific gravity of a mineral?
2. How is it determined?
3. A piece of lead, whose specific gravity is 11.4, weighs in the air, 456 grains; what is the weight of an equal bulk of water?
4. How much does this same piece of lead weigh in water?
5. A piece of glass weighs in air 24 ounces; in water 16 ounces: what is its specific gravity?
6. What is "vitreous lustre?"
7. What is the "streak" of a mineral?
8. What is the melting-point of ice?
9. Describe the expansion and contraction of water.
10. At what temperature does water reach its maximum density?
11. Describe an experiment to prove this.
12. Explain the structure of ice.

CHAPTER III.

A GLASS OF MELTED ICE.

Is water melted ice—or is ice frozen water? In one sense both statements are true, but the latter is the popular form, while the former is the more accurate and scientific. It is more accurate because ice is the natural or normal form of this mineral when left to itself, and it is only by the application of heat that it is forced into the form of water and kept in that form; it is more scientific also because it agrees better with our way of thinking of other mineral substances, the great majority of which are most familiar to us in their "ice," or frozen, or solid state. Thus when we speak of iron, lead, copper, or sulphur, we naturally think of the solid metal or stone, and when these substances are melted we do not give them new names as in the case of ice and water but speak of them as melted iron, melted sulphur, etc. At all events there is a gain for our present purpose in thinking and speaking of the mineral which we are studying as ice, and in regarding what is usually called a glass of water as a glass of melted ice. There may be people living in the frigid zone to whom this matter of thinking has always appeared the natural, perhaps the only one.

We are now ready for a simple experiment with a glass of melted ice.

Hold a lump of sugar so that one corner of it dips into the melted ice, and first observe the sugar. It will be seen that, beginning at the bottom it rapidly changes from white to transparent gray, and that it is growing wet and beginning to soften and dissolve. It seems to absorb the water like a sponge, and in a short time is evidently soaked through. This rise of the liquid into and through the sugar is explained in the same way as the rise of oil in a lamp-wick, or through any very small tube; that is to say, as due to the drawing power of the sides of the minute openings through which the liquid rises, added to the drawing power or attraction of the particles of liquid for one another. The attraction between the solid and liquid is called *adhesion*, because the liquid adheres to the solid; the attraction between the particles of the liquid is called *cohesion*, because the particles cohere to one another. The walls of the tubes or pores which contain the liquid attract those liquid particles which are nearest them and thus cause a thin film of the liquid to spread out and upward on their surfaces. As this film of the liquid is forced upward it draws with it the whole column of liquid, and this continues up to the point when the weight of the rising column equals the attractive force. From this it will be seen that this attraction is not a new force, but simply the united action of the ordinary forces of adhesion and cohesion; and it is also evident that the elevation of the liquid will be greatest in the smallest tubes, since the weight of the column to be raised is there least in proportion

to the amount of liquid which comes in contact with the walls of the tube. This explanation was first given by Leslie in 1802. The Latin word for hair is *capillus*, and since the phenomenon I have been describing is best observed in very small, hair-like or "capillary" tubes, this form of force is called "capillary attraction."

Let us now repeat the experiment—this time, however, turning our attention to the melted ice under the sugar. As soon as the sugar touches the liquid fine wavy and syrup-like streams are seen to flow from it down into the glass, where they presently grow thinner and disappear. This continues until the last particle of sugar crumbles in the hand and loses itself in the liquid below. The sugar is now said to be dissolved. Substances which can be dissolved are called "soluble," and the liquid which has the power of dissolving them is called a "solvent."

Find by experiment as many different substances as you can that dissolve in water, and ascertain whether heating the water renders their solution more or less easy and rapid. In connection with these experiments test the temperature of the water before and after the addition of the substance to be dissolved.

Make a strong solution of alum and allow the water to evaporate slowly. It will be observed that a portion of the alum is deposited from the solution in the form of crystals; and the important fact is learned that crystallization occurs not only on the solidification of minerals by freezing—as in the case of ice—but that it

also usually accompanies the solidification of a dissolved mineral on the evaporation of its solvent.

Professor Tyler suggests the following interesting experiments. Make a strong solution of blue vitriol, which is a compound of sulphuric acid and copper. Put this in a warm place and suspend a pebble or cinder in it. Notice the form and color of the crystals which form upon the pebble. Now mix equal parts of alum and blue vitriol, make a strong solution of the mixture and allow the water to evaporate. The resulting crystals are most interesting. There is a mixture of the two solids, but the crystals of the alum are distinct from those of the vitriol, and with care you can almost completely separate the blue crystals from the white. Thus the minute particles which have been held in solution not only combine in crystal form, but particles of the same substance unite in crystallization with one another, and not with those of another substance even when mingled in the same solution.

It appears also from this experiment, that solids in solution do not necessarily unite with particles of their solvent so as to form a new and distinct substance, but that they may be merely mixed with those particles, though in so finely divided a state as to be indistinguishable to the eye.

Water has been called the "universal solvent," not only because of its universal distribution, but because it dissolves a larger number of substances, both solid liquid, and gaseous, than any other mineral.

This solvent power of water is one of its most val-

uable properties. On it depend many of the phenomena of nature, and many of the processes of art and manufacture. The substances that dissolve in water are so numerous, and so universally distributed, that it is doubtful whether a single gallon of chemically pure water could be collected in the whole world from natural sources. Of these soluble substances salt is one of the most abundant, and countless rivers are constantly carrying one or both of its elements to the sea.

Professor Tyler says, that although the proportion of dissolved matter in the ocean is not large, yet it is always growing greater, and it has been estimated that there is now enough in the whole ocean, if it could be separated, to form a mountain range larger than the Alps.

The water that issues from springs and wells is usually "hard;" that is, impregnated with lime or some other mineral in solution, which decomposes soap, and thus renders the water unfit for washing. Even rain-water, which is the purest form in nature, is tainted with impurities, caught and dissolved by the water as it falls through the air.

To this property of water we owe all our mineral springs; to it we must refer our deposits of iron ore. It is the cause of the formation of great caves in the earth with their wonderful and exquisite stalactites and stalagmites; it is even largely responsible for the valleys between the mountains, and for much of the scenery of the world, which it reduces to forms, smiling and peaceful, or frowning, abrupt, and grand. With

out it the vegetable world could not draw its support from the earth, the blood could not circulate in the veins of animals, and all life would perish from the earth. Do not pass this subject hastily. Think about it. Make a list of as many other advantages and disadvantages that arise from the solvent power of water as you can.

Do not forget its cleansing power on the one hand, nor on the other the dangers that arise from its power of dissolving poisonous or putrefying substances.

In connection with this last suggestion, consider the wisdom of removing as far as possible from your water-supply every thing that is unclean or unwholesome. Typhoid fever and many other dread diseases result from drinking water which holds in solution or in suspension the invisible but certain seeds of death.

Owing to the selective power of the crystallizing particles of any mineral, by virtue of which they unite only with others of the same sort, as already shown in the case of alum and blue vitriol, ice freely formed upon a lake is more free from impurities than the water under it. There is a truth in the popular saying, that the "dirt freezes out." Yet it is a dangerous error to suppose that all ice is pure, and safe for domestic use. If dirty water freezes solid, all the impurities contained in the water will be retained in the ice. They do not unite in crystallization with the particles of ice, but they are caught, and held in the ice, just as a stick would be if plunged in freezing water; and when the ice is melted they are again set free, and the freezing does not destroy their unwholesome and dangerous

properties. Even when only the surface of foul water freezes, these impurities are but partially expelled, so that it is a safe rule that dirty water makes dirty ice, and deadly water makes deadly ice.

Let us now return to our glass of melted ice. You remember that this was produced from a transparent solid by the action of heat, and that it is kept in its present liquid form by the continued action of the same force. We will now make the experiment of applying a greater amount of heat to the melted ice.

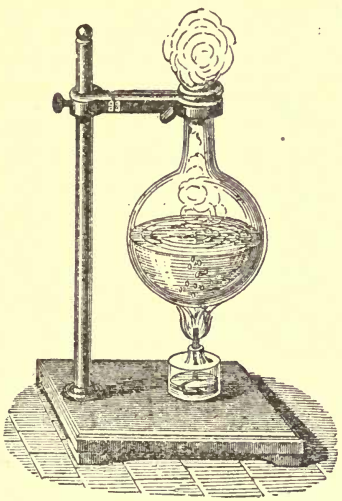

Fig. 9.

Pour the water into a glass flask, or a tin cup, put a thermometer in it, and place it over a spirit-lamp, or gas-jet, or set it on a stove. Fig. 9. At first no change is observed, except that by close observation you may see that the liquid is expanding; but as the temperature approaches 212 degrees Fahrenheit, bubbles appear on the bottom and sides of the vessel where it is hottest. These first bubbles are due to the separation of particles of air which have been previously dissolved or absorbed by the water. A little later, small bubbles begin to come

from the heated portion of the water, and as they rise through the cooler liquid above they disappear. The formation and disappearance of these bubbles is accompanied by a peculiar "singing" sound. After this, larger bubbles rise and burst on the surface, and little clouds of vapor begin to appear, while the whole mass of water is agitated. The thermometer now stands at 212 degrees Fahrenheit. This phenomenon is called boiling, or, in scientific phrase, *ebullition.*

As this boiling continues, the water is observed to be growing rapidly less in quantity, and after a short time wholly disappears. The thermometer meanwhile has remained stationary at 212 degrees. Now, nothing has been seen to pass from your flask except the vaporous clouds already mentioned, and it thus becomes evident that unless some other way can be discovered to account both for the disappearance of the water, and also for the appearance of the vapor, we must conclude that the water has been changed into the vapor, and has passed out of the vessel in that form.

It is astonishing that so large a quantity of water can so rapidly escape in so airy and intangible a form; and the vapor itself has so completely disappeared in the air that we find it hard to believe that the water which has been "boiled away" has not actually been destroyed. If not, where is it? In order to settle this point, we must devise some method of catching, and examining the vapor, instead of allowing it to escape.

Hold a cold glass jar, mouth down, over the neck of the flask. As it becomes filled with vapor you will

observe that its sides grow foggy, and presently drops of water appear, and begin to trickle down inside.

Is it not plain that the vapor is only another form of water, forced into this condition, and kept there by heat? So much of this heat now leaves the vapor and is expended in warming the cold sides of the jar, that there is not enough left to hold the water in its vaporous form, and it again becomes liquid, and re-appears in the form of water. This return of vapor to the liquid form on the withdrawal of heat is called *condensation*. The process of boiling a liquid to transform it into vapor, and then cooling the vapor and collecting again the condensing liquid, is called *distillation*, and as the liquid on vaporizing leaves most of its impurities behind, this is the simplest and best method of obtaining most liquids, and water among them, in the purest possible condition.

You may wonder why I have not called this water-vapor *steam*. It is not steam in the strict sense. Steam is invisible. If you look at the neck of your flask while the water is briskly boiling in it, you will see no appearance of vapor, and the flask above the water appears quite empty; yet it is really filled with steam. On leaving the mouth of the flask this steam becomes partially condensed in the air, and then appears in vaporous clouds. Indeed, I have here used the word "vapor" in the popular rather than the scientific sense. Strictly speaking, the invisible steam is the true vapor, and the white cloud is composed of minute particles of water condensed from the vapor.

Another phenomenon presented by water bears a close resemblance to the production of steam by boiling, yet with important differences.

From all exposed surfaces of water there is a constant and usually unnoticed passage of a portion of the water into the form of vapor. This is called evaporation. It differs from boiling in two respects, *first*, it takes place only at the surface, and slowly; while boiling produces steam throughout the mass of the liquid, and rapidly; *second*, it occurs at no fixed temperature; while boiling, as we have seen, occurs, under ordinary pressure, only at 212 degrees Fahrenheit. Vapor is formed directly, even from ice, without intermediate melting. Ice "wastes away," even in cold weather; and our laundry-women are right in saying that clothes hung out in winter "freeze dry." It would take many volumes to trace the results of boiling and evaporation. Modern civilization depends largely upon steam, which not only cooks much of our food, but furnishes most of our power. It would seem as if the old Arabian tales of Aladdin's lamp, and the Fisherman and the Genie, were intended as parables to illustrate the wonders of steam. Aladdin rubs his lamp, the emblem of fire, and the strange fire-spirit appears, ready to do his bidding. From the Fisherman's casket arises a huge mass of vapor, which gradually assumes the form of a puissant being capable of any act of helpful ministration or mad destruction.

To-day with our lamp we have evoked a spirit no less powerful; capable like the Genie of expanding

until the sky seems full, and of contracting to the limits of an iron pot; when under due control, it drives our locomotives across the continent, runs our looms and spindles, heats our homes and lights our streets; if uncontrolled, it tears our engines to pieces, wrecks our habitations, and destroys our lives. Like electricity, magnetism, and the explosive force of powder, steam is one of the invisible but mighty brotherhood of fire-spirits, and repeats to-day to all who know how to keep it in control, the formula used by the obsequious Genie of old, "I, and the others, slaves of the lamp!"

From every rivulet and river, from every lake and sea, from the whole expanse of ocean, vapor is constantly rising, invisibly or in fantastic wreaths of mist; and this collects in spreading clouds which fill the sky, cover the world with cooling shade, and gently fall in showers that refresh the thirsty earth, bringing life to every herb and tree, to animals and man; or else, uniting in hostile fury, they sweep over the earth in a cyclone; then "heavily they fall on the sea, and from its very bottom crash down the whole expanse."

QUESTIONS ON CHAPTER 3.

1. What keeps water in its liquid form?
2. Give an illustration and an explanation of "capillary attraction."
3. How is the temperature of water affected when salt is dissolved in it?
4. Why is water called the "universal solvent?"

5. What important results follow this solvent action of water?

6. Does foul water "freeze clean?"

7. Describe the process of ebullition.

8. At what temperature does water boil?

9. Distinguish between steam and other vapor.

10. Describe and explain "condensation."

CHAPTER IV.

IS WATER AN ELEMENT?

From the experiments and observations outlined in the preceding chapters, we have learned that water exists in three different forms or "states," namely, as a solid, a liquid, and a vapor, or gas. It has also been learned that these forms follow one another in the order given upon the addition of heat, and that upon the gradual removal of the heat the steam again condenses into water, and the water crystallizes into ice.

Whether other minerals under like conditions pass through these same three states must be left for your future experiments to decide; though you know by former experience that different substances, such as lead, iron, sulphur, and wax may exist both in the solid and liquid form; and you know that others, as alcohol, naptha, petroleum, and ether, are easily changed from liquids to gases.

Consider that expression "different substances." Are ice and water and steam three different substances? We have seen that they are only three different forms of the same substance; and a proof is that one can be converted into the other at will.

Suppose, now, in the first of three cups you have a quantity of leaden bullets, in the second a number of

iron balls, and in the third a mixture of iron balls and leaden bullets: have you in the cups three substances or two? If you stir a tablespoonful of powdered chalk into a glass of water, you obtain a milky fluid unlike either. Is it a new substance or only a mixture of the two? Dissolve a teaspoonful of salt in a glass of water. The salt disappears, and we seem to have only one substance, namely, the water, left. Have we really one or two? Is brine a new substance, or a mixture of two?

Simple as these questions appear they nevertheless have an important bearing upon our study. In the case of the leaden bullets and iron balls, we evidently have no new substance, for we can easily separate the two. In the case of the chalk and water the decision is not so immediate, but if we allow the mixture to stand, the particles of chalk settle at the bottom, and the two substances separate themselves. The brine may seem more puzzling, but if we evaporate the water we can recover all the salt. The main difference between the three cases is that the particles of the substances mixed are united more closely in one case than in another. In every instance the test is to find whether we can separate the mixture into the simple substances, or "*elements*," of which it is composed.

Most of the things we see are compounds. A substance so simple that it cannot be separated into other substances, nor be shown to be made of other substances, is called an element. Chemists recognize at present between sixty and seventy elements; but the number is

not settled by any means. It has often happened that a substance which has for many years been considered elementary has been found, after all, capable of separation, or "analysis," into simpler substances; while every little while new substances are discovered which have to be added to the list of elements until, and unless, they can be proved to be compounds.

On the whole, as our power of analysis increases, the tendency is to reduce rather than enlarge the number of elements; and there are some who are even asking whether, after all, those ancient philosophers may not have been right, who held that all the forms of matter were derivable from some one, though undetermined, element.

From all this you will perceive that *analysis*, or the separation of mineral compounds into their elements, is one of the most important means of learning their true character. In evident mixtures, like that of iron balls and lead bullets, the analysis consists merely in the act of picking out one kind, and leaving the other. In the case of solids held in liquids but not dissolved, the analysis often results by the process of settling; and in the case of solutions by evaporation or distillation. Certain compounds of iron can be ana'yzed by heat. A mixture of iron-filings and sand can be analyzed by holding a magnet near it; the iron is drawn to the magnet, and the sand left behind.

Some substances are resolved into their elements by a current of electricity.

A few hundred years ago water was believed to be an

element. "Fire, water, earth, and air," were said to be "the four elements," of which the whole earth is composed.

Let us now make a few experiments with a view to learning whether this ancient opinion of water is correct.

So far as we can perceive by the most careful examination of pure water under ordinary conditions, it is a simple substance. The eye detects no mixture, and nothing is revealed either to taste or smell.

No matter how long we may allow it to stand, there is no separation into simpler substances either by settling or evaporation. If we boil it and condense its vapor, we get again the same quantity of the same substance, pure water, and nothing else. The approach of a magnet produces no perceptible effect.

Let us try electricity. The following experiment is more difficult than any I have before described, but with a little care and patience you can perform it successfully, and the result is so surprising, so beautiful, and so instructive, that you cannot afford to neglect it.

Before attempting it, however, we must learn some of the simpler facts about this wonderful form of force.

Cut out a piece of sheet zinc as large as a silver half-dollar. Place this under your tongue, and place the silver half-dollar on your tongue. Now bring the forward edges of the zinc and silver together.

Do you feel anything? This peculiar tingling sensation is produced by a current of electricity, which passes through your tongue from the zinc to the silver,

and back again to the zinc, at the point where the two metals touch each other in front. This course of the electricity is called an electric circuit, and such a circulating current is set in motion whenever two different metals are placed in any acid, and joined together outside the acid, either by being brought into contact directly, or by means of connecting wires. To illustrate this more clearly, fill a glass or earthen jar with dilute sulphuric acid, and immerse in the liquid at a little distance from each other a plate of zinc and a plate of copper. Attach a wire to the upper end of each plate, fig. 10.

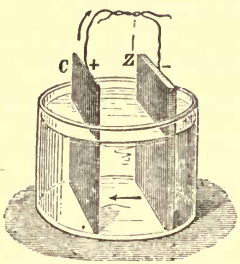

Fig. 10.

Now, taking a wire in each hand, touch one to the upper and the other to the under surface of your tongue. Is not the result similar to that produced by the coin and the piece of zinc, though more intense? Bring the ends of the wires together. What is the effect?

This apparatus is often called a "cell." Many other substances besides zinc and copper produce similar currents. One of the most common and practical substitutes for copper is *carbon*, a substance with which you are familiar in the forms of charcoal and coke, and the hard incrustations which accumulate in gas-retorts. This is not the place, however, to enter into a description of the numerous varieties of cells which have been devised. They all work on the same principle, and can be

bought at a low price. The combination of two or more cells is called a "battery." For our present experiment you will need a battery of four or five cells; and, if you do not wish to make them, you can probably get them at second hand from any telephone-exchange at trifling cost. The free ends of the connecting wires, through which the force of the electricity manifests itself, are called *electrodes*, a Greek word, meaning "electric pathways."

These ends are also called "poles," a word which means in this connection the points of greatest intensity or tension.

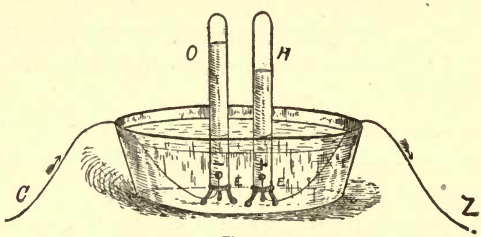

Fig. 11.

Having now provided your electric battery, fill a glass vessel with water, to which, in order to assist the passage of the electric current, some sulphuric acid must be added, as pure water is an imperfect conductor. Fasten a piece of platinum to the end of each wire; or in other words, let the connecting wires end in platinum electrodes. This is not essential, but like the sulphuric acid greatly facilitates the action of the electricity. Put the electrodes in the water, and invert over each a glass

tube filled with water, fig. 11, in which O and H are the tubes, C the wire coming from the copper or carbon, Z the wire coming from the zinc, and E E, the electrodes.

On the passage of a current of electricity, bubbles are seen to rise from each electrode. The water gradually falls in the tubes, and that, too, without perceptibly rising in the large vessel. It looks as if the current were boiling the water and filling the tubes with steam. But, look again. *The effect in the two tubes is not the same.* In one the apparently empty space above the water is twice as great as in the other; and yet the water in that tube does not seem to "boil" any faster than that in the other.

At all events we are rapidly losing sight of our water. Let us carefully examine the apparently empty tubes to ascertain whether it can be hiding there.

We will first determine whether they are full of steam. Placing the hand on them we find them cold; this proves that steam is not present, and indeed a little reflection would have convinced us of this, from the fact that the tubes were neither burst nor forced out of the water, nor clouded with condensing vapor.

Are they empty then?

Light a match or taper, and blow it out, leaving the end aglow. Remove the tube O from the water, first placing your finger over the opening of the tube, and holding it closed until you have turned its mouth upward, and thrust the glowing—not blazing—end of the taper into it. The glowing taper suddenly bursts into

flame, and burns vividly. In fact this tube is filled with a remarkable sort of gas, which we shall soon study more in detail. For the present it is enough to know that it has been separated from the water by electricity at the electrode which is connected with the copper in the battery; that it causes other substances to burn very brilliantly, and that its name is *Oxygen*.

Now remove the other tube, holding its mouth downward, and apply a lighted match. This tube, too, is evidently filled with gas, but of a different sort, for it catches fire at the mouth of the tube, and burns with a flame which is pale and bluish, but very hot. This gas also we shall study later, contenting ourselves just now with the knowledge that the electricity sets it free from the water at the electrode which is attached to the zinc, that there appears to be twice as much of it as there was of the oxygen, that it burns with a pale, hot flame, and that its name is *Hydrogen;* a Greek word, which means "a producer of water."

If the current of electricity were continued long enough, all the water would be separated into these two gases, which, moreover, always exist in the water in a constant proportion, namely, one volume of oxygen to two volumes of hydrogen.* In order to prove the correctness of this analysis, we must ascertain whether we can produce water again by recombining the two gases. If they are mixed in the proper proportions, they form at first a merely mechanical mixture like

* Recent experiments by Professor Morley show that water contains 2.0002 parts of oxygen to 1 of hydrogen.

that of coal gas and common air; but the moment a flame is applied, a much closer union is brought about —in fact, a *chemical combination* occurs—accompanied by a violent explosion, and in place of the two commingling gases we have—water.

This experiment must be made only with a small quantity of the gases, such as can be contained in a test-tube, for on a larger scale it is dangerous.

It is possible, however, to cause oxygen and hydrogen to combine quietly, by bringing them in tubes from separate receptacles into a common burner. The grandest experiment of this nature, according to Professor Huxley, was made by three eminent French chemists. It continued from the 13th to the 22d of May, 1790. During this time the apparatus was constantly watched, the experimenters sleeping alternately on mattresses in the laboratory; 25,964 cubic inches of hydrogen, and 12,571 of oxygen were consumed, and the union of these gases produced 7,244 grains of water.

It is a startling thought that all the water on the globe has been formed at some time or other by the union of oxygen and hydrogen, which in their free state are known only in the physical condition of gases.

QUESTIONS ON CHAPTER IV.

1. In what three states does water exist?
2. Find by experiment two other substances that resemble water in this.
3. What is an element?

4. What is analysis? Give an illustration.
5. Describe an electric cell; a battery.
6. What are electrodes? Poles?
7. What is added to water before testing the effect of electricity upon it? Why?
8. Describe the effect of an electric current on water.
9. What two gases are obtained from water?
10. In what proportions are they combined to form water?
11. How can they be distinguished?
12. How can they be combined?

CHAPTER V.

FIRE.

In the previous chapter, we learned by our experiments that it would be quite wrong to regard water as an element, for by a current of electricity we analyzed it into two gases, oxygen and hydrogen.

We also learned that at a high temperature these two gases unite with a hot flame, and return again to the form of water.

These are among the most beautiful and significant phenomena in nature; and, with those we have before studied, give us a key to the intelligent understanding of the whole world of matter. We have been led to the discovery of two of the most important elements in the world; we have learned the nature of an element; and we can now understand that every material object consists either of such an element alone, or in a state of simple mixture with other elements, or in a more intimate union with other elements in the form of a chemical compound. We have seen that mixtures can be mechanically separated; and that, in the case of water, even the bands of chemical union can be loosed by the force of electricity; and we have learned, finally, that elements can be caused to unite chemically, and that, in the case of oxygen and hydrogen, this union is accom-

panied by light and heat. Such a combination is termed *combustion*, or burning.

Whenever you see fire, therefore, you may be certain that a rapid combination of at least two elementary substances is taking place. From this you will perceive that the old notion that fire is an element was also incorrect.

Let us now continue our experiments with a view to learning whether oxygen will combine with other substances than hydrogen. Tip the end of a bundle of fine iron wire with burning sulphur, and plunge it into a jar of oxygen. Do not let the surprising brilliance of the result distract your thought from the lesson. The most brilliant teachers are not always the best. Turn your attention from the strange sight of the blazing iron, to the dull results of the bright combustion. Notice that when the fire ceases the oxygen is gone and the iron is gone: they are "burned up." Are they then destroyed? No more than the water was destroyed which we "boiled away." The dark, cindery-looking stuff in the jar is the hiding-place of both; rather it *is* both, chemically combined.

That you may be convinced that this is true, you should repeat the experiment more carefully, taking such precautions as will assure you that none of the oxygen escapes from the jar during the combustion, and that nothing else gets in except the oxygen and the iron, and the small bit of sulphur used to heat the iron. You must also determine the exact weight of the oxygen and of the iron consumed, and of the substance

left after their combustion. If you do this accurately, and if you find as other experimenters have found that the weight of the iron burned added to the weight of the oxygen used up is exactly equal to the weight of the substance left after combustion, you can hardly doubt that the latter is formed by the union of the oxygen and the iron. Substances formed by combustion with oxygen are called *oxides*; and this compound of oxygen and iron is called *iron oxide*.

The heat of combustion is the result of the chemical combination, and the light is due to the heat. The heat is caused by the friction of the little particles of the combining substances as they are forcibly drawn together; just as an anvil is heated by hard blows from a hammer. The more rapid the combustion the greater the heat. Oxygen and iron are combining slowly at ordinary temperatures whenever iron grows rusty; and the iron-rust is an oxide of iron.

I said you should weigh your oxygen. Let me show you how. You need a glass or globe holding three or four quarts, the neck of which is closed by an air-tight stop-cock, and adapted to be screwed to the plate of an air-pump. The air is then pumped out, and the weight of the empty globe ascertained by a delicate balance. Oxygen is now allowed to enter, and the globe again weighed. The difference in weight will be the weight of that quantity of oxygen; and knowing this, you can calculate the weight of any other quantity

It may save you trouble, however, to know that a pint of oxygen weighs $11\frac{3}{10}$ grains, and a cubic foo

1¾ ounces. (*Faraday*). Oxygen is nearly sixteen times as heavy, bulk for bulk, as hydrogen, which is the lightest element known; so that, although, as you remember, we obtained twice as much hydrogen from water, yet the oxygen was nearly eight times as heavy as the hydrogen; in other words, if we reckon by weight, instead of bulk, oxygen makes up $\frac{8}{9}$ of water, and hydrogen only $\frac{1}{9}$.

I am glad to come back in this way to water, for I think this is just the right time for you to make another experiment with it; one that you could not have understood so clearly before you saw the oxygen and iron uniting by combustion, and forming the black oxide of iron.

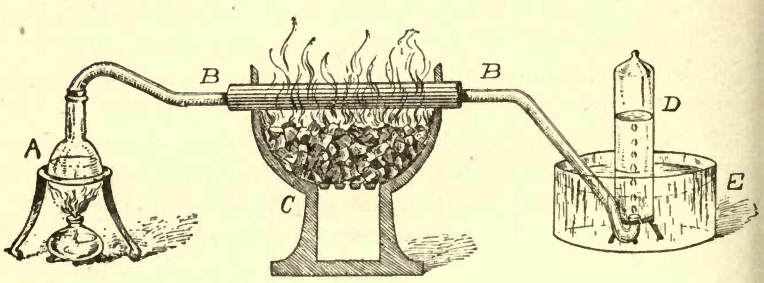

Fig. 12.

For the experiment you will need apparatus arranged practically as shown in Fig. 12. A is a flask for boiling water over an alcohol lamp; C a small furnace; B B, a tube about an inch in diameter passing through the furnace; D, a jar of water inverted in a vessel of water, E. The tube BB may be a gun-barrel, a piece

of gas-pipe, or preferably a porcelain tube, and is filled with iron turnings or fine bright iron wire. The heat of the lamp boils the water and converts it into steam. The steam passes through the red-hot tube, and bubbles arising through the water are caught in the upper part of the jar D. When all the water in this tube has been expelled, you will know that its place has been taken by some sort of gas. Is it steam? No, for it is cold. Testing it with a match you find that it catches fire and burns at the mouth of the tube with a pale blue flame. You recognize hydrogen; and you perceive that in passing through the furnace, the steam has parted with its oxygen. Recollecting the experiment of burning iron with oxygen, you now readily understand that what you saw going on so brilliantly in the jar has taken place unobserved in the red-hot tube; namely, the oxygen of the water has combined with the iron in the tube, and the hydrogen has passed on alone into the jar D.

When the tube BB becomes cool, you will find in it the same dark-colored substance that you found when the oxygen and iron combined in the jar; namely, iron oxide.

This experiment is commonly made to illustrate a method of obtaining hydrogen, but we have used it to impress more firmly upon our minds the manner in which a chemical union takes place between oxygen and heated iron; and to illustrate the new fact that the tendency to this combination is so strong that it not only occurs when heated iron is plunged into pure oxygen,

but that the oxygen in water will separate from the hydrogen in order to unite with the iron.

It is an important fact that a stronger attraction exists between particles of heated iron and oxygen, than between particles of oxygen and hydrogen. A like principle controls the combination of all the other elements; they manifest apparent preferences in their combinations, somewhat as young people do in their choice of partners; and by learning what these preferences are, and taking advantage of them we can effect many innocent divorces, such as that we have just caused between hydrogen and oxygen, and bring about many not unhappy marriages like the one just effected between oxygen and iron.

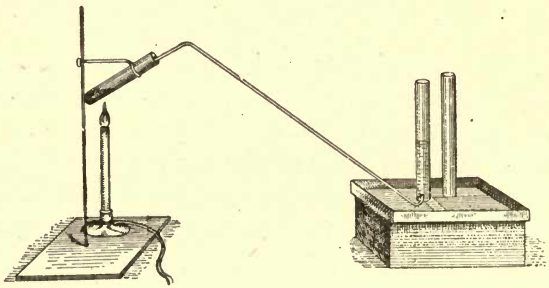

Fig. 13.

This special attraction existing between certain elements is called *chemical affinity*.

Substances which combine at certain temperatures may often be separated again by increasing the heat. Thus the iron oxide which we have obtained can be so

intensely heated as to be resolved again into its elements; but so great heat is required that the experiment is difficult. A similar compound of mercury and oxygen, however, known as oxide of mercury, can be very readily separated by heat; and affords us a practical method of obtaining oxygen for our experiments.

For this purpose apparatus should be arranged as in Fig. 13, which needs no explanation.

The most convenient and rapid method of preparing oxygen, when a considerable quantity is required, is to heat a mixture of equal parts of potassium chlorate and black oxide of manganese.

Potassium chlorate contains three elements, oxygen, chlorine, and potassium, two of which are as yet strange to you. Oxide of manganese, as its name implies, is a compound of oxygen with another unfamiliar element, manganese. We do not care to study these new elements at present; it is sufficient to know that oxygen is combined with them, and can readily be disengaged by heating.

Experiment.—Mix an ounce of potassium chlorate with an equal weight of coarsely powdered oxide of manganese in a mortar. Heat a little of the mixture in a test-tube before using, because the oxide of manganese is sometimes impure, and may then cause explosions if mixed with potassium chlorate and heated. If the substances separate quietly in the tube, you may safely proceed with the experiment on a larger scale, by heating the mixture in a glass retort as shown in

Fig. 14, and collecting the oxygen over water as usual. The retort, however, frequently cracks.

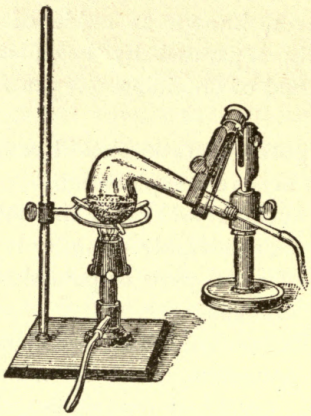

Fig. 14.

Having thus collected as much oxygen as you wish, exercise your ingenuity in trying all sorts of experiments with it. So long as you do not mix hydrogen with it, it is quite safe. Smell it, taste it, breathe it, burn all sorts of substances in it (by the way it will combine with almost every thing in the world), blow soap-bubbles with it, and see whether they rise or fall; put bits of cold iron in it, and leave them for a few days to see whether any combination takes place without heat; in a word, get acquainted with oxygen, so that you can recognize it wherever you meet it, and so that you can recognize its effects even when you do not detect its presence.

It will take some time, but you can well afford it, just as I can well afford to give large space to it in this little book. It is the most important element in the world. It is eight-ninths of all our water, it constitutes one-half the solid crust of the earth, it makes up four-fifths of the weight of all our herbs and shrubs and trees, and three-fourths of your own body, and of the bodies of all animals, and it is the chief source of fire. We walk on it, and burn it, and eat it, and drink it, and breathe—

Ah! but I did not mean to tell you just yet that one-fifth of all the air is oxygen. I meant to let you find it out by degrees. But, never mind, it is true, and we must devote the next chapter to a study of the air. Before passing to that, however, I think you are entitled to know an easier method of obtaining a supply of hydrogen. You should prepare enough to enable you to become as familiar with it as with oxygen.

The most convenient process depends upon two facts: first, that acids contain hydrogen; second, that, as a rule, when a metal is placed in an acid, all of the acid except the hydrogen enters into new combinations in which the metal is taken up, leaving the hydrogen free.

Metals are the successful rivals of hydrogen. No matter how well satisfied oxygen and sulphur and chlorine and the rest may appear to be with hydrogen for a partner, once let almost any metal come among them, and they fly to it, leaving the hydrogen deserted.

To illustrate this fickleness of the elements, and at the same time to procure a good supply of sufficiently pure hydrogen, try the effect of a little zinc upon that compound of sulphur, oxygen, and hydrogen, known as sulphuric acid.

Provide a jar fitted with a cork having two holes. Through one hole pass a funnel-tube, and through the other a glass tube bent in convenient form. Put a small handful of granulated zinc, or zinc clippings, into the jar.

Prepare a mixture of sulphuric acid and water as follows: Into six ounces of cold water pour very slowly one ounce of ordinary sulphuric acid, keeping the mixture constantly stirred. If you pour it in rapidly the heat evolved may be so great as to convert the water into steam and cause the strong acid to spatter. After the mixture has cooled, pour enough of it through the funnel into the jar to cover the zinc.

A brisk ebullition occurs at once, so eager are the little particles of oxygen and sulphur to combine with the zinc, and the deserted hydrogen rises into the jar. For two or three minutes the jar will be filled with a mixture of hydrogen and air; and since one-fifth of the air is oxygen, this is a highly explosive and dangerous mixture, as we have already learned.

Wait a few minutes, therefore, until the hydrogen has expelled all the air, and then collect as much as you wish in the usual way over water. Fig. 17.

You can easily ascertain whether all the air has been expelled by filling a test-tube with the escaping gas,

and observing whether it burns quietly when lighted. In this, as in all other experiments with hydrogen, remember to hold the jar containing it mouth downward; otherwise it will either escape or become mixed with air, and explode when ignited. You had better collect several bottles of hydrogen, for there are many interesting and instructive experiments to be made with it. Invent as many as you can. Burn a fine jet of it in air under a bright dry glass, and notice how the glass

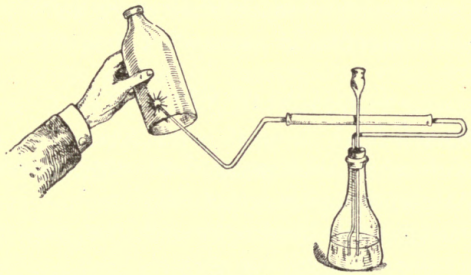

Fig. 15.

at once becomes clouded by small drops of water condensed upon its surface. You readily understand from this experiment that the hydrogen is combining, or burning, with the oxygen of the air. Push a burning taper up into a jar of hydrogen. It is extinguished. Blow soap-bubbles with it, and you have beautiful balloons. These can be blown by connecting a tobacco-pipe by a rubber tube to the tube from which hydrogen is driven by the action of the acid and zinc. Dip the bowl of the pipe in soap-suds, and the force of the escaping gas should be sufficient to blow bubbles.

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Finally blow some bubbles with a mixture of air and hydrogen, and touch them with a lighted taper as they rise, taking care, however, not to light them while on the pipe, because the explosion might force flame into the jar and blow it to pieces.

Thus the oxygen and hydrogen separated by the intruding metal are again united, and in these most fairy-like, and many-colored balloons are floating away on their bridal tour. A touch of flame bursts their crystal car, but serves only to render their union more intimate, and they fall together in a sparkling drop of water.

QUESTIONS ON CHAPTER 5.

1. What is fire?
2. With what substances does oxygen combine? Ans. All the elements except fluorine.
3. Describe the combustion of iron and oxygen. What is the product?
4. How can a gas be weighed?
5. What are the effects of passing steam through a red-hot tube filled with iron-borings?
6. What is chemical affinity?
7. Describe a practical method of obtaining oxygen.
8. Has oxygen any effect on cold iron?
9. Is it heavier or lighter than air?
10. Where is it found in nature?
11. Is it in the air?
12. Describe and explain the process of obtaining hydrogen by the action of zinc upon sulphuric acid.
13. What practical use is made of hydrogen?

CHAPTER VI.

AIR.

If you have carefully studied the preceding chapter, and performed the experiments there described, you have become pretty familiar with two of the most important elements. The most characteristic property of oxygen is its power of combustion with other substances. You can hardly have failed to notice that the burning of a candle, the flame of a gas-jet, and the fire on the hearth all resemble the flame of substances burning with oxygen, differing only in intensity.

A simple test of the presence of oxygen in the air is holding a cold glass jar or a piece of any cold metal over the flame of an alcohol lamp, or over a gas flame. Both alcohol and coal gas contain a large percentage of hydrogen.

Presently you will observe a dampness coming upon the cold glass or metal, and this increases until it stands in drops and trickles down. On examination, you will find this liquid to be pure water. Now water cannot be produced except by combining oxygen and hydrogen. In this case the hydrogen comes from the alcohol or gas; and it is plain that the oxygen must come from the air.

But if air is so largely composed of oxygen, how is it

that the air does not cause the alcohol and the gas to burn as brilliantly as oxygen does? And why will not iron burn in the air? And if oxygen unites in combustion with everything else, why is it that, once a fire gets started, particularly a great conflagration as of a forest or a city, it does not increase and spread and burn up the whole world? Well, I have no doubt it would do that if there were nothing but pure oxygen

Fig. 16.

in the air. Nor would it wait for a great fire to start it, for there are substances like phosphorus that combine with oxygen at ordinary temperatures.

If the air were pure oxygen, such substances would kindle first, and they would serve as a match to set the world on fire. Iron buildings would blaze more fiercely than wooden ones. That would be a grand holiday for the fire-spirits!

"The elements would melt with fervent heat, and the heavens would be rolled together like a scroll."

This terrible result is prevented only by the controlling presence of another element in the air, which holds the fiery oxygen in check, and compels it to combine as a rule in a reasonably temperate and safe combustion.

Let us see whether we can unmingle a little air and procure some of this wonderful gas which alone saves the world from bursting into flames.

Float a light porcelain dish on water in a glass or earthen basin. Put a small piece of phosphorus in the floating dish, and light it. Cover the dish with a large jar.

The phosphorus at once combines with the oxygen of the air, forming a phosphorus oxide, at first appearing in the form of a white cloud, which, however, soon dissolves. When the jar which has become heated by the burning of the oxygen and phosphorus becomes cool, you will notice that one-fifth of that part of the jar that was before filled with air is filled with water, which has risen in it from the basin.

You may now test the contents of the upper four-fifths of the jar. It cannot be empty, or the weight of the atmosphere would press the water from the basin up into it and fill it completely. It cannot be steam, for it is cold. It cannot be hydrogen, for it will not burn. It is not oxygen, for a lighted taper thrust into it, instead of blazing more brightly, is immediately extinguished. It has no color, taste, or smell. It is completely inactive. No action short of the most intense electric force can cause it to combine chemically

with oxygen. It is a perfectly safe substance, under ordinary conditions; and its name is *Nitrogen*.

But this quietness is no indication of lack of power. "Deep rivers run still." It takes a man of considerable strength of character to live quietly among noisy neighbors, particularly if he succeeds in keeping them moderately quiet too. So, lest you should feel any contempt for this most reserved and indifferent gas, I must warn you that on due occasion it enters into combination

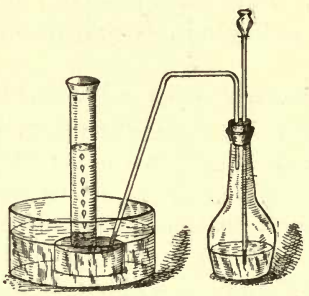

Fig. 17.

with other elements, and manifests the most sudden and unexpected activity. Thus when under the influence of electricity, or by the decay of plants and animals, nitrogen is forced to unite with hydrogen, it forms ammonia and in this form possesses a pungency and caustic flavor that quite compensate for its ordinary lack of smell and taste. Pure ammonia is a colorless gas which is very soluble in water. The "liquor ammoniæ" of the drug-stores is a solution of this gas in water, prepared as shown in Fig. 17, by passing

the gas into a flask containing water and kept cool by being placed in a large vessel of cold water. This is necessary, because considerable heat is caused by the dissolving gas.

When nitrogen is combined with a gas called chlorine it forms one of the most dangerous and explosive compounds known.

Dulong, the chemist who discovered it in 1811, lost one eye and three fingers for his pains, and two years later Faraday and Davy although on their guard met a similar accident. "Knowing that the liquid would go off on the slightest provocation, the experimenters wore masks of glass, but this did not save them from injury. In one case, Faraday was holding a small tube containing a few grains of it between his finger and thumb, and brought a piece of warm cement near it, when he was suddenly stunned, and on recovering consciousness, found himself standing with his hand in the same position, but torn by the shattered tube, and the glass of his mask even cut by the projected fragments. Nor was it easy to say when the compound could be relied on, for it seemed very capricious; for instance, one day it rose quickly in vapor in a tube exhausted by the air-pump, but on the next day, subjected to the same treatment, it exploded with a fearful noise, and injuring Sir Humphrey Davy."

The name, nitrogen, signifies a "producer of niter," and is applied to this gas because it is one of the principal elements in niter, or "saltpeter." Niter is formed by the combination of nitrogen with oxygen and potassium.

It is found in long, colorless, six-sided crystals, marked by fine parallel lines or "*striæ*." Minerals so marked are called "striated." It has a strong salty taste, and can readily be dissolved in water. It occurs as a natural product in the East Indies, Egypt, and Persia. It is found in the juices of various plants, as the sunflower nettle, tobacco, and barley.

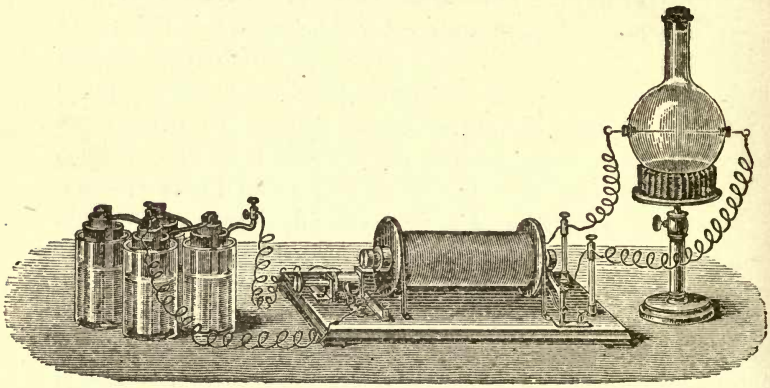

Fig. 18.

Niter is used in the manufacture of sulphuric and nitric acids, as an ingredient of fireworks, and especially in the manufacture of gunpowder.

It is extensively used in medicine, particularly in acute cases of rheumatism, in certain affections of the throat, and in spasmodic asthma.

Nitric acid is an important compound of nitrogen, oxygen, and hydrogen, and in one form, in which water also is present, is known as *aqua-fortis*. It is astonish-

ing that these two elements, which when simply mixed form the air we breathe, when chemically combined, make an acid so powerful and destructive.

If a large glass globe filled with dry air be connected with an electric battery and an "induction coil," as shown in Fig. 18, and sparks be passed through the air in the globe, red fumes are rapidly formed. On pouring a few drops of water into the globe and shaking it, the red fumes are absorbed by the water, and nitric acid is the result.

When a flash of lightning passes through the air the same effect is produced, and this accounts for the presence of a certain amount of nitric acid in the atmosphere.

When nitric acid is exposed to the air it gives off strong fumes which are dangerous to breathe. The pure acid absorbs moisture very rapidly. It is extremely corrosive; for example, when dropped on the hand it produces painful wounds, and colors the skin yellow or brown.

Nitric acid is used in the formation of coal-tar colors, in the manufacture of gun-cotton, nitro-glycerine, and dynamite, and in nitrate of silver, which is so widely used by photographers.

This acid has a powerful effect upon sulphur, phosphorus, carbon, tin, and many other substances, the oxygen contained in it uniting rapidly with these elements and "oxydizing" them. Pour a little nitric acid upon granulated tin and study the result. The white powder which is deposited is an oxide of tin.

When cotton-wool is soaked in a mixture of strong

nitric acid and sulphuric acid, washed, and dried, it becomes "gun-cotton," an explosive so powerful, that when used in shells it is six times as effective as the same weight of gunpowder. No smoke accompanies its explosion, and for this reason it has been very useful in mining. Other advantages are that time, dampness, and exposure to air do not injure it; it occupies a comparatively small space, is light, and is not liable, as gunpowder is, to accidents from spilling. Dynamite and other more recently invented compounds have, however, now largely taken the place of gun-cotton.

Another curious compound of nitrogen and oxygen must be mentioned, differing from nitric acid in the proportions of the two gases, and in the absence of hydrogen. It is called "nitrous oxide," and is remarkable for the effects it produces when breathed.

These are first, "singing" in the ears, then insensibility, and then, if the breathing be continued, death. When one volume of oxygen is mixed with four volumes of nitrous oxide, the mixture is called "laughing gas," and the effect of inhaling it is less serious. It causes temporary insensibility during which minor surgical operations, such as the extraction of teeth, may be performed without pain. The effect soon passes off and does no harm. Owing to the liability of having impurities mingled with this gas, it is dangerous for amateurs to experiment with it.

These are but a few of the many interesting compounds of nitrogen, and are given to illustrate how an element which, when pure, appears to be the nearest

possible approach to nothing at all, having no color, taste, or smell, being incapable of burning, or supporting combustion, and unable to support life, becomes, when combined with other elements, suddenly endowed with the most intense energy; helping to produce the most vivid colors, the most acid taste, the most pungent smell, the most violent explosions, and the most necessary constituents of food.

To illustrate the importance of nitrogen in food, it may be stated that it occurs abundantly in all flesh, eggs, and cheese, and is an essential ingredient of all foods directly adapted to produce blood and muscle.

I have refrained from giving minute directions for obtaining the various compounds of nitrogen, as their preparation and use involve too much danger; but if you have attentively performed the few simple experiments given, you must by this time have become reasonably familiar with this most fascinating element.

How largely does the beauty, the power, the life, the very existence of this world, depend upon the constant but unobserved activity of three invisible and intangible substances, oxygen, hydrogen, and nitrogen!

QUESTIONS ON CHAPTER 6.

1. How can you prove the presence of oxygen in the air?
2. What prevents the immediate destruction of the world by fire?
3. How can the elements mingled in air be separated?

4. Describe nitrogen.
5. What is ammonia?
6. What is the meaning of the word, nitrogen?
7. What are the uses of niter?
8. How may nitric acid be produced?
9. What are its effects and uses?
10. What is gun-cotton? Its advantages?
11. What is laughing gas?
12. The value of nitrogen in food?

CHAPTER VII.

EARTH.

We have learned that the old notion that water, air, and fire are elementary substances, is incorrect. It is hardly necessary to add that those who held this opinion were equally wrong in thinking that "earth" was a fourth and final element, if by "earth" they meant what is now commonly understood, namely, the soil under our feet, for I have already told you that about one-half of this consists of oxygen in combination with other substances, and that hydrogen and nitrogen are also found in it in vast quantities.

Before going on to study the form of earth known as soil, it is only fair to the ancient philosophers from whose writings this popular theory of the "four elements" took its rise, to say that they never meant anything of the sort. They were quite intelligent gentlemen, and knew as well as we know, that there are more than four different substances in nature; or at all events more than four apparently different substances. What they intended to say was more nearly this, that all the elements of the world exist in four different states or conditions; namely, solid, liquid, gaseous, and a fourth state still further removed from the solid form—a state in which matter still more rarefied surrounds and interpene-

trates all the grosser forms. This closely resembles that subtle and almost hypothetical form of matter which modern philosophers believe to pervade the earth, and the remote spaces beyond the air, and which they call "the ether." I do not say that the ancients had as clear and scientific a knowledge of the stuff the earth is made of as modern students have, but the words, "solid," "liquid," "gas," and "ether," are much more accurate interpretations of their classic formulæ, than the popular but misleading words, "earth," "water," "air," and "fire."

Let us now examine a handful of earth from the garden, not attempting to discover, perhaps, all its elements, but carefully enough to satisfy ourselves that it is composed of several different substances.

I have here a little earth in a saucer. It is a dark-brown substance, resembling, but for color, some of the cheaper grades of brown sugar. There are little grains in it, with here and there small lumps which crumble easily at a touch. Bits of straw and tiny root-fibers can be seen without close observation. These evidently come from the fertilizer which was spread over the garden last fall, and plowed in this spring.

I will now put some of the earth under a microscope. Using at first a hand-lens of low power, it seems that this earth is composed of tiny grains of sand of different colors. Some are black and dull, some white, some transparent and glassy, some brown, some thin, scale-like, and glistening. It is a curious mixture. I now place a smaller quantity on a glass slide, and examine it

under a more powerful lens. It is at once transformed from a brown powder into a heap of tiny pebbles of various colors, shapes, and sizes. Most of these tiny pebbles are transparent, and of a brownish tint; many of them have a crystalline appearance. My garden is clearly composed mainly of a fine sand or loam. The tiny pebbles under the glass seem to be mingled with a finer and smoother substance which I recognize as clay; and the presence of minute particles of vegetable matter is everywhere more visible than to the naked eye.

Let me now heat some of the earth. I put it in this test-tube, and hold it over the flame of an alcohol lamp. If I had no test-tube I should put it on a shovel, and cover it with an old tumbler.

The first thing noticed as the earth grows hot is a change of color. It is becoming lighter—at the same time the sides of the tube are growing foggy, and now drops of water are trickling down in it. A quantity of water was in the soil, and it has been converted into vapor and driven off. After longer heating the earth becomes quite dry, and now there appear here and there little sparks and small glowing coals, that flicker for a moment and go out. The bits of straw and roots are burning up. First they grow very hot, and give off little puffs of smoke and gas; then they turn black like bits of charcoal; then when still hotter they combine with oxygen of the air with a bright red glow, and then they crumble down into gray ashes. I now put the earth in the test-tube between two pieces of glass, press them tightly together, and slip the upper one

along over the earth. I hear a grating sound, and on examining the glass, I find that it is covered with distinct scratches.

So I might go on and make many other interesting experiments with this handful of garden-soil, but enough has been done for our present purpose. We have learned beyond question that earth is not an element, but a mixture of many different substances.

We have found oxygen and hydrogen in the water driven off, carbon in the charred bits of straw and roots, clay binding together the little grains of sand, and these sand-grains themselves of various sorts, some soft and smooth, others hard enough to scratch a pane of glass. It is plain enough that these bits of charcoal are not essentially different from the larger pieces on the hearth; that these minute particles of clay are just like the large masses which we find in clay-beds; and that these tiny pebbles which constitute the sand are different only in size from the larger pebbles found in general. It will be more convenient to study them in their larger forms and separately, and so we may now throw away our earth, and examine the substances of which it is composed as we find them in larger specimens. We will begin with the hard stone that scratches glass.

In the collection accompanying this book you will find a specimen of it, No. 18. It is a piece of quartz. Observe first its physical properties, and I shall ask you to determine for yourself its degree of hardness, specific gravity, lustre, color, and streak, and its fusibility. Test it also to discover whether it is brittle or

not. When you break it for this purpose, observe whether it splits evenly as a piece of slate does, or whether the surfaces after breaking are irregular. The manner in which a mineral breaks is often important as indicating its structure, and therefore as a step toward identifying it. The peculiar breaking of a mineral is called its *fracture*. If it splits more or less evenly, such splitting is called *cleavage*, instead of fracture.

Drop a little hydrochloric acid from the end of a glass rod upon the specimen. What effect does it have?

Specimen No. 19 is a crystal of quartz. You will observe that it is six-sided, or hexagonal, and if it has not been broken, it is also finished at one or both ends with an hexagonal pyramid. These crystals may be found in almost every part of the world, for a large part of the solid crust of the earth is composed of quartz. They are often found loosely scattered in sand, which is formed by crumbled or "disintegrated" rock, but they naturally occur in veins which run through other sorts of rock.

Quartz crystals were believed by the ancients to be merely a kind of ice frozen too hard to melt; but they are now known to result from a solution of the mineral in water, which trickling through crevices in the rock gradually evaporates, and slowly builds up the crystals from the quartz dissolved in it. The name "crystal," a Greek word for ice, is still retained as a result of the old and erroneous opinion. Nothing can be more delightful than the finding of a crystal vein; it repays many days of searching. Some years ago one of my pupils

in Lenox Academy discovered a few loose crystals of quartz remarkable for their brilliancy and symmetrical form. It was suggested that as these had been found in a regular line along the crumbling edge of a rock, a little digging might reveal a vein; and the boys were invited to meet in the field at a certain hour prepared to "stake out their claims," and enter upon a small mining operation. The appointed hour brought a most amusing scene. One boy had a large hammer and a drill; another a shovel and an old iron spoon; another a crowbar and a garden trowel; another a post-hole digger that opened and shut like a huge pair of shears; while one excited youth appeared with a can of powder and a fuse. Little labor, however, was needed. A few strokes of the shovel removed a layer of earth, and a cry of pleasure arose from all the group. A narrow crevice in the limestone rock was revealed, and it was filled from end to end with gleaming crystals of transparent quartz; some loose, others united in groups and clusters of great beauty; and ranging from the size of a pea to that of a butternut. Some were tinged with yellow, others clear as water. Some were nearly regular in form, but most of them were curiously distorted, yet every one was hexagonal, every angle measured the same number of degrees as every other, and each had on one end, at least, a pointed pyramid.

Down dropped the little miners on hands and knees, regardless of torn trousers and finger-nails. When the sun went down all were still at work, with their earth-stained heads all down out of sight, and their heels all

up in the air; and every boy learned more about quartz in that short summer afternoon than he could have dug out of books in a month. So can you if you will follow their example, and make a careful examination of the country within ten miles of your own home. You may not find a vein of crystals, but you will find plenty of interesting mineral specimens of some sort to repay you for your pains. You should make a mineralogical map of your town, upon which you will indicate the location of the various minerals you find; and you should make a collection of all local minerals, labelling each specimen carefully, and being particular to add the exact place from which it was taken.

Possibly one reason that quartz so closely resembles ice is that more than half of it is composed of the principal element of water; namely, oxygen.

The other element in quartz is new to you; it is called *silicon*, and is never found in nature except in combination with oxygen, from which it is, however, extremely difficult to separate it. In spite of this difficulty, however, it is so important for you to get acquainted with the element, since next to oxygen it is the most abundant in the world, that I will tell you how the chemist Berzelius obtained pure silicon in 1823.

He took advantage of the well-known preference of fluorine for metallic over non-metallic substances, on the same principle by which we obtained hydrogen when we passed steam over red-hot iron. You remember that in that case the oxygen left the hydrogen and united with the iron.

Berzelius put ten parts of dry potassium silicon fluoride—a compound of the metal potassium with silicon, and a gas called fluorine—into an iron tube; his object being to induce the fluorine and potassium to leave the silicon: he then added nine parts of metallic potassium, and heated the mixture red-hot. This produced a violent commotion among the elements in the tube, and when it was all over, and the tube cooled down, he found, as he had hoped, that the charms of the metallic potassium had proved strong enough to overcome even the attachment which had previously existed between the fluorine and the silicon. The fluorine and potassium were united together in a new compound—*potassium fluoride*—and the deserted silicon was left alone.

Before Berzelius got hold of it, however, he had to shake up the whole mixture, first in cold and then in hot water, to dissolve out the potassium fluoride. When this was done, he found the pure silicon settled at the bottom, and on taking it out and drying it, he had a brown powder, and was perhaps the first human being to see and handle this wonderful and most important element. I do not know why you may not succeed in repeating his experiment and sharing his delight.

On experimenting with this silicon powder, it was found that neither sulphuric nor nitric acid affected it; but when heated in the air it readily entered into combustion with oxygen, forming an artificial quartz, which often melted around particles of the silicon powder, leaving a portion of it unburned in the centre.

When minerals occur in an uncrystallized state, as

when carbon is in the form of charcoal, or silicon in the form of this brown powder, they are called "amorphous"—which simply means without definite form.

By other processes silicon has been obtained in dark and glittering eight-sided crystals, and also in regular double six-sided pyramids, of a dark steel-gray color. It is not a metal.

QUESTIONS ON CHAPTER 7.

1. Is "earth" an element?
2. What did the ancient philosophers mean by speaking of the "four elements, fire, water, earth, and air?"
3. Describe the results of your examination of a handful of garden soil.
4. What substances are most abundant in it?
5. Describe the physical properties of quartz.
6. Of what two elements is quartz composed?
7. What then is its chemical name? (Oxide of silicon.)
8. Describe a crystal of quartz.
9. How are quartz crystals formed?
10. Describe silicon.

CHAPTER VIII.

QUARTZ.

In the previous chapter, by an examination of a handful of loam we learned that earth is not one element, but a mixture of many minerals, chief among which is quartz. We further learned the more prominent physical properties of quartz, and its chemical composition.

How large masses of quartz rock are split by frost, broken down and rounded into pebbles by the action of weather and water, and at last reduced to fine sand, will be more fully explained in the third volume of this series, where we hope to treat somewhat in detail of the forces which are constantly at work, changing the form and modifying the structure of the earth's crust; but as this mineral constitutes a very large part of this crust. we shall be justified in devoting a few more pages to it now.

No mineral appears in more numerous and varied forms than quartz. We have seen it in its transparent forms as a fragment of rock, and as a gleaming crystal. In these forms it rivals in transparency the finest glass. From it spectacles are sometimes made, and for this purpose Brazil furnishes some of the finest pebbles; but it is doubtful whether it posesses any real advantage over glass, except that being much harder it is less

likely to be scratched. The ancients made exquisite goblets, cups, vases, and seals of rock-crystal. Mr. Ruskin pronounces a perfect sphere of rock-crystal one of the most beautiful objects in the world. These are made in Japan, sometimes five or six inches in diameter and costing hundreds of dollars. In their perfect clearness and perfect shape they rival the transient glory of soap-bubbles, while their hardness and exquisite polish render them almost indestructible.

One of the common varieties of quartz is flint. It is from *silex*, the Latin name for flint, that the name *silicon* is derived. Flint varies in color from nearly black to light brown, red, yellow, and grayish white, and is sometimes veined, clouded, marbled, or spotted. It breaks with a curved fracture, called "conchoidal," or shell-like, and is translucent. Its coloring is due to the presence in minute quantities of other substances, such as lime, iron, and carbon. It is abundantly found in beds of chalk. Its property of striking fire with steel is well known, and before the invention of friction-matches led to the general use of flint for kindling fires and firing guns. Old "flint-lock" muskets and pistols are still preserved in historical museums.

Sparks may be struck by steel from any hard mineral, but flint sparks are so hot and brilliant that it is thought that they are due not only to particles of the stone heated to redness by friction, but that a chemical combination of *silica* (*silicon dioxide*) and iron takes place.

The most ancient use of flint was probably for sharp weapons and cutting-tools. Flint knives, axes, and

arrow-heads are among the most interesting relics of the Indians and other savage peoples. The principal use of flint at present is in the manufacture of fine earthenware, and particularly in the preparation of the delicate glaze or enamel which covers, protects, and adorns it.

An interesting natural form of quartz is the *geode*. Geodes are approximately round shells of rock, lined with crystals, or with minerals in distinct layers. They range in diameter from a quarter of an inch to a foot or more. When water holding a mineral in solution finds its way into a hollow, and deposits the mineral there in layers or in crystals, a geode is formed. If the mineral in solution is silica, quartz crystals may result. If the quartz is deposited in layers, these usually appear with different shades of color, and such variegated layers of quartz are known as agate. If the color of the quartz remains uniform, and particularly if it is deposited in curved forms like clusters of grapes, and of a waxy lustre, it is called chalcedony. If the color is red, we have a geode of carnelian, sard, or sardius.

The question naturally arises, how the cavities in which these crystals and layers occur are formed. One way is by the production of steam inside the rock material while that is in a soft or "plastic" state; as is seen in the sponge-like slag or *scoria* of a furnace or a volcano. Rocks which were once in a melted condition, like the "trap-rock" of the Palisades along the Hudson river, afford some of the most beautiful geodes. Another method by which the cavities may be made, is

by the dissolving action of water on fossils. The solvent water eats out the fossil shell or coral.

Perhaps a clergyman might find a sermon in these stones. The hidden crystals must remain entirely lustreless until the geode is broken. Though they have the power of shining so brilliantly, they come to their glory only when shattered. In losing their life they find it.

Onyx and sardonyx are varieties of agate, in the first of which the alternate layers are white and dark brown or black; and in the second, reddish yellow or orange.

Moss-agates, or mocha-stones, are formed by the infusion of a solution of manganese or some other mineral into the quartz, where it often closely imitates the branchings of moss or little fibers of root. It is possible that these forms may in rare instances be due to actual fibers or sprays of moss, which were once imbedded in the quartz, and dissolved out by the water; in which case they may be considered as tiny geodes.

Prase, or chrysoprase, is a beautiful pea-green variety of quartz, and is often found in alternate layers with white chalcedony.

One of the handsomest varieties of quartz is amethyst, which by the oxide of manganese or iron is tinged with brilliant purple. The name "amethyst" means "a preventer of intoxication," and was given to this mineral by the Greeks, who believed that from a cup of amethyst one could drink wine and escape a headache. Fine specimens are found in Lincoln County, North Carolina, on the shores of Lake Superior, and in Colorado. The finest varieties are from Brazil, Ceylon, India, and Siberia.

The inspired St. John is largely indebted to quartz for his conception of the glory of the New Jerusalem: "The building of the wall thereof was jasper: The first foundation was jasper; the third chalcedony; the fifth sardonyx; the sixth sardius; the tenth chrysoprase; the twelfth amethyst." Surely a mineral which not only composes much of the crust of the present earth but which is also to be found in such abundance in the foundations of the "New Earth" that is to be, deserves our careful study!

Another beautiful compound of silicon must be mentioned here, which differs in composition from quartz only in containing from 5 to 13 per cent. of water, although, as with quartz, iron, and other substances frequently give it a tinge of color. This is opal. It is never found crystallized. It has a conchoidal or shell-like fracture like chalcedony, and is very brittle. The finest variety is called "precious" or "noble," or "oriental" opal. It is translucent, of a bluish or yellowish hue, and exhibits a beautiful play of rainbow tints. The finest opals come from Hungary. Common opal is less clear, and does not have any play of color. Wood-opal is a petrifaction retaining the form and apparent structure of the wood which has been dissolved away, and replaced by the silicious mineral.

Three beautiful varieties of common rock-crystal remain: milky-quartz, which is opaque and white as snow; rose-quartz, which is transparent and of the most delicate rose-color; and smoky-quartz, which is translucent and tinged with grayish brown. It will be well to

notice that all varieties of quartz fall naturally into two groups: (1) the clear, glassy kinds, like rock-crystal, amethyst, rose-quartz, and the like, and (2) the waxy or dull kinds, such as chalcedony, agate, and flint.

Quartz constitutes nearly all the sand of the sea-shore, and of the great deserts. The small particles of most other minerals are dissolved by water, while those of quartz remain hard and "sharp."

Of the numerous uses of sand, I mention five. You are perfectly familiar with its use on sand-paper, and as a covering for "silicate" slates, and blackboards.

In 1870, Benjamin Tilghman, of Philadelphia, invented the "sand-blast," a rapid stream of sharp sand driven against glass, stone, or metal, for the purpose of cutting, boring, or engraving. By its use, also, blocks of stone may be turned in a lathe into circular and other forms. Pilasters have in this way been finished in a few hours, while if cut by hand as many days would have been needed.

Sand is one of the chief ingredients in plaster and mortar, being united with lime to form a strong and durable cement.

Its chief use, however, is in the manufacture of glass, which is composed of quartz, united with lime, soda, metallic oxides, or other substances.

There is a legend that glass was accidentally discovered ages ago, by certain Phœnician merchants, who were returning home in a ship laden with soda, and who were compelled by a storm to land on a sandy shore under Mt. Carmel. They propped up their cook-

ing-pots by lumps of soda placed on the sand, and when these and the sand were melted by their fires the first glass was formed; but this is not the place to enter into a history or description of glass-making, nor even to enumerate the most important of its countless uses.

Although quartz is so useful in the construction of our dwellings in the form of sand for mortar, it is no less valuable when naturally compacted and mingled in various forms of building-stone. All our sandstone, for example, is largely composed of quartz, and quartz is one of the necessary ingredients of granite, gneiss, and many other rocks.

Finally, quartz or silica occurs in many vegetable structures, giving them strength and hardness, and often, as in the case of stalks of grain, a fine enameled surface. It also occurs in sponges in little needle like "spicules."

Consider the incalculable evil that would result from the sudden destruction of the silicon that is in the world! The contents of every glass tumbler and jar and bottle, would be let loose; microscopes and telescopes and spectroscopes would vanish. Rain and wind would have free access to our homes through empty window frames; stone and brick buildings would crumble to the ground; mountains and the solid crust of the earth would sink away, and the oxygen now safely held in the strong grasp of silicon would be let loose to consume the whole world with fire.

In closing this chapter, I must remind you that in using the expression "crust of the earth," I do not mean the whole mass of the earth, but only an outer shell, a few

miles thick; no thicker in proportion to the whole earth than the varnish on a school-globe. Geologists know nothing by personal observation about the materials of which the great bulk of the earth is composed. They have only scratched around a little here and there on top. There are reasons known to astronomers, however, and others understood by electricians, which make it not incredible that the earth is a solid, or nearly solid, mass of iron, covered over with a layer or a few thin layers of various sorts of rock, and topped with a sprinkling of pulverized rock, called earth or soil, which is altogether too thin and scanty to be taken largely into account.

QUESTIONS ON CHAPTER 8.

1. Describe the appearance and uses of rock-crystal.
2. Derivation and meaning of the words crystal and silicon?
3. Describe flint, and tell its uses.
4. What are geodes, and how are they formed?
5. Describe agate, sardonyx, and onyx.
6. How are "moss-agates" formed?
7. Describe prase, amethyst, and opal.
8. Uses of sand?
9. If you have seen glass manufactured, describe the process.
10. What results would follow the sudden removal of silicon from the earth?
11. What is meant by the "crust" of the earth?

CHAPTER IX.

A LESSON IN CHEMISTRY.

It is evident from what we have now learned that changes are constantly taking place in mineral substances. We have seen ice changing to water, and water to steam. We have seen iron bursting into brilliant flame in oxygen, and changing into iron oxide. We have seen water changing into two gases, and these gases changing back again into water. We have seen various solutions changing into vapor and crystals. Thinking over all these changes you will see that they may be divided into two classes:

1. Those that do not alter the composition of substances.

2. Those that alter the composition of substances.

Changes of the first kind are called *physical* changes. Those of the second kind are called *chemical* changes.

The change from water to ice is a physical change.

The change from water to oxygen and hydrogen, or the reverse, is a chemical change.

From the experiments we have made, we have also learned that chemical changes may be caused: (1). By heat, or electricity. (2). By bringing different substances into contact. (3). By bringing different substances together in a solution.

In all cases of chemical action, we have observed that the substances acted upon lose their distinctive properties, and that new substances with different properties are formed.

Notice now that all cases of chemical action may be divided into three classes: (1) Composition; (2) Decomposition; (3) Double decomposition.

Double decomposition occurs when substances act upon one another and give rise to two or more new substances instead of one. A good example of double decomposition was shown in the process by which Berzelius separated silicon from potassium silico-fluoride, page 83.

What it is that causes substances to combine with one another is not known. To the unknown cause of their combination, the name *chemical attraction* has been given. We must now study carefully the proportions in which the different elements combine.

Two facts have been learned by experiment:

1. *The same chemical compound always contains the same elements in the same proportion by weight.*

Water, for example, always contains the same elements, oxygen and hydrogen, and in the same proportion by weight, namely, 16 to 2. This is called the *law of definite proportions.**

2. *The proportionate weight of each element in every compound is always expressed by a particular number, or by a multiple of that number.*

* Recent experiments by Professor Edward Morley indicate that the exact proportionate weight of oxygen in water is between 15.877 and 15.882.

For example, the proportionate weight of oxygen in every compound is always expressed by 16, or by some multiple of 16.

This is the law of multiple proportions.

These facts can be explained by supposing each element to be made up of little particles of uniform weight.

For illustration, if you have a number of iron balls, each weighing seven ounces, and a number of glass balls, each weighing three ounces, it is plain that, however you may mix the iron balls with the glass balls, the total weight of glass in the mixture can always be expressed by three or some multiple of three, and the total weight of iron by seven or some multiple of seven

Now, since all elements combine just as they would if they were made up of particles of uniform weight, chemists assume that they really are made up of such particles; and since these particles are the smallest possible parts into which an element can be divided, they are named *atoms*, a Greek word meaning *indivisible.*

It is believed that these tiny atoms do not absolutely touch one another, and that they are never absolutely at rest. If a mass of gold, for example, could be viewed through lenses of sufficient power to reveal its structure, we should probably see as wild, and apparently confused, a dance of atoms as that of the motes in the sunbeam, or the planets in the sky. I say "apparently" confused, for there is really no disorder either among the atoms, or the motes, or the planets; but each moves in accordance with a regular and well defined law.

Finally, it is thought these atoms occur in groups of two (or in some elements more, but a constant number for each element), and that these groups are never broken up except when the element enters into combination with some other substance. These little groups of atoms are called *molecules*, a word meaning simply "little masses." In the giddy dance of atoms, the molecules move together like inseparable partners in a waltz.

Molecules of a single element may be called *simple molecules*. It is evident that the smallest particle, or molecule, of a compound substance must be made up of atoms of the different elements of which it is composed. A molecule of water, for instance, is made up of atoms of oxygen and hydrogen. Such molecules may be called *compound molecules*. The weight of a molecule is equal to the sum of the weights of the atoms that compose it.

The rapid motion of atoms and molecules to which I have referred is closely connected with the phenomena of heat. The hotter they are, the more rapid and extended are their movements; and conversely, the more rapidly the molecules are made to move, the hotter they grow.

It would hardly be correct to say that heat makes them move, or that their motion causes heat; it would perhaps be better to say that their motion *is* their heat. At all events, as this motion or heat, or both, increases, the substance expands, just as it would if it were made up of little masses which were separating farther and farther

from one another and getting away from the power of one another's attraction. As the heat increases the solid passes into the liquid, and the liquid into the gaseous form. In the gaseous state the molecules are no longer held together, but are violently driven farther and farther asunder.

Avogadro's Law: In 1811, a chemist named Avogadro announced as the result of many experiments that *equal volumes of all gases under the same conditions of temperature and pressure contain the same number of molecules.*

This is known as Avogadro's law. If this be true, then by weighing equal volumes of gaseous substances we can learn the relative weights of the molecules of these substances.

From the molecular weights thus determined, the atomic weights of the elements have been learned.

As hydrogen enters into combination in smaller proportion by weight than any other element, its atomic weight is taken as the unit. When we say, therefore, that the atomic weight of oxygen is sixteen, we mean simply that one atom of oxygen is sixteen times as heavy as one atom of hydrogen.

Chemical symbols.—As the chemist is mainly occupied in studying the combination of atoms, it has been found convenient to represent these atoms by symbols, and the most convenient symbols are the letters which stand for the names of the elements. Thus the letter O is used to represent one atom of oxygen, the letter H, to represent one atom of hydrogen, etc. When the

names of several elements begin with the same letters, as in the case of indium, iodine, and iridium, a second letter is added to distinguish them; thus In for indium, Ir for iridium; and when this is not practical, as in the case of iron, where I or Io might be taken for iodine, Ir for iridium, and In for indium, the chief letters of their Latin names are used; thus, Fe, from the Latin "ferrum," iron, is the symbol for that element.

The student should now carefully examine the table of elements and symbols on page 263. It is not best to commit them to memory at once. They will gradually become familiar by use. The numbers after the symbols are the atomic or combining weights of the elements. When we desire to express the fact that various elements are put together, we write their symbols one after the other with the sign (+) plus between them. Thus 2H+O means that two atoms of hydrogen are added to one atom of oxygen. If we wish to express by the symbols that the elements thus brought together have united in chemical combination, we omit the sign + and write the symbols closely together as follows: H_2O. Such an expression is called a *chemical formula.* Chemical formulas have a puzzling look to the beginner, but they are really of the greatest use both in simplifying our language and in rendering chemical combinations easy to understand.

When the result of bringing certain elements together is their chemical union, we express that fact in the following form: $2H+O=H_2O$; which is read: Two atoms of hydrogen added to one atom of oxygen

combine to form water. The sign of equality (=) is very convenient here. Such a formula is called a *chemical equation.*

As the symbols of the elements represent atoms, it is evident that the symbols of compounds represent molecules. The symbol for water, H_2O, represents, according to the atomic and molecular theories already explained, the smallest particle of water that can exist as such. It is made up of two atoms of hydrogen and one atom of oxygen, chemically combined.

Exercise.—Read the following equations:

$Hg+O=HgO$ (*Mercuric oxide.*)
$Na+Cl=NaCl$ (*Sodium chloride* or common salt.)
$2N+2O=N_2O_2$ (*Nitric oxide.*)

It will be observed that before combination takes place the number of atoms in any element is represented by a large figure written on a line with the letters; while the number of combined atoms is represented by a small figure written to the right and below the symbol to which it belongs.

In laboratory work combinations are less frequently made between two elements directly than between compounds, and it is in expressing completed combinations, or "reactions" as they are called, that the value of symbols is most apparent.

Recur for a moment to the experiment by which Berzelius obtained pure silicon, page 83. Observe how many words were needed to explain the combinations he effected, and see how much simpler it is to express the same reaction, thus:

K_2SiF_6	+	$4K$	=	$6KF$	+	Si
(*Potassium-silicon fluoride.*)		(*Potassium.*)		(*Potassium fluoride.*)		(*Silicon.*)

The symbols tell the whole story in one line; not only giving us the substances brought together and the compounds that result, but also the exact proportions of each by weight.

A little study of this equation will show you that the weight of the resulting substances is precisely equal to the sum of the weights of all the elements that are combined. Turn to the table of atomic weights.

The *atomic* weight of K being 39.04, K_2 = 78.08
The *atomic* weight of Si being 28, Si = 28.00
The *atomic* weight of F being 19.10, F_6 = 114.60
The *atomic* weight of K being 39.04, 4 K = 156.16
And the sum is..... 376.84

Now add the weights represented in the other side of the equation.

The weight of K being 39.04, and the weight of F being 19.10, K F = 58.14; and 6 K F = 348.84
The weight of Si being 28, Si= 28.00
And the sum is again..... 376.84

This is true of every chemical equation, and forcibly illustrates a fundamental principle of the science, namely, that *nothing is ever destroyed by any chemical process.*

Water seems to boil away, but every molecule of it continues to exist in the form of vapor. This vapor may be decomposed into its elemental gases, but every atom that formerly existed in the water is still pre-

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served in the gas. Iron was burned up in oxygen, but both were found in their full weight in the iron oxide that remained.

As no element is ever diminished or destroyed, so no elementary substance ever increases in total quantity. When coal is burned in a furnace, the weight of the vapor and smoke that escape from the chimney is about four times as great as that of the coal consumed, but this is only because in the process of combustion nearly three tons of oxygen have combined with every ton of carbon.

QUESTIONS ON CHAPTER 9.

1. Define chemical and physical changes.
2. State three methods of producing chemical changes.
3. What are the three classes of chemical changes?
4. What causes chemical combination?
5. Give the laws of definite and multiple proportions.
6. What hypothesis explains them?
7. Distinguish between atoms and molecules.
8. What is Avogadro's law?
9. What is the unit of atomic weight?
10. What is the advantage of chemical symbols?
11. Can the total quantity of any element be increased or diminished?
12. How does this appear in chemical equations?

CHAPTER X.

A PIECE OF CHARCOAL.

Having gone out of our way in the last chapter to consider some of the laws that govern the elements in their chemical combinations, and to become acquainted with the symbols used to express those combinations, or "reactions," let us now return to the examination of our handful of earth.

You remember that, besides a quantity of water, we had already found in it a large number of minute grains of quartz, composed of a very interesting element, *silicon*, in combination with oxygen. Notice, in passing, that as a rule, each mineral has two names, one given and used by mineralogists, as in this instance *quartz;* the other used by the chemist to indicate its composition, as *silicon oxide*, or abbreviated, SiO_2.

We next observed bits of charcoal scattered throughout the soil. These are identical in substance with the large pieces of charcoal on the hearth, and consist of the element known as *carbon*.

Carbon is the chief element in the composition of all the substances which we burn for the sake of heat or light; including wood, coal, gas, oil, tallow, and wax. Whenever these fuels are heated to redness in a vessel from which air is excluded, their other elements, hydrogen

and oxygen, are first driven off in gaseous forms, partly as water and partly in combination with carbon, and a large portion of the carbon remains behind, mingled with a little ash. If, now, free access of oxygen be allowed, the carbon combines with it gradually, not with a flame but with a steady glow, and the product of this combustion is an invisible gas, which we shall presently study.

The purest form of charcoal is lamp-black. Charcoal does not crystallize, and is therefore an amorphous form of carbon.

You are so familiar with the physical properties of charcoal that you need not devote much time to them now. The specific gravity of this substance, however, is greater than at first appears. A piece of charcoal floats on water, and appears thus to have a specific gravity of less than 1; but this is due to its porous condition. It is buoyed up by the air which fills its pores. If it be finely powdered it sinks in water, and its specific gravity is found to range from 1.5 to 2.

Coke is a form of carbon made by heating coal without free access of air, as is done on a large scale in the manufacture of illuminating gas. Coke bears about the same relation to coal that charcoal bears to wood.

Lamp-black is a very finely divided form of charcoal, which is deposited on cool objects placed in the flames of burning oils. Most artificial lights deposit a black soot on objects placed in them.

This soot is used in the manufacture of printer's ink.

Carbon is acted upon by very few substances, and is

nearly insoluble, so that it is impossible to destroy the color of printer's ink without destroying the material upon which it is impressed.

Bone-black, or animal charcoal, is made by charring bones or other animal substances.

Bone-black and wood charcoal have the power of absorbing gases, and are therefore useful as purifiers of foul air. Charcoal also absorbs many of the impurities contained in foul water, and is extensively used in *filters*. A charcoal filter should be of good size, and the charcoal should be occasionally renewed.

Some coloring matters are removed from liquids by passing them through filters of bone-black. This is of practical advantage in the refining of sugar. The syrup first obtained from the cane or beet is strongly colored, but the color is removed by filtering through bone-black.

Many impure varieties of amorphous carbon are included under the name *coal*. They all bear a more or less close relation to charcoal, and, like it, have been formed by the gradual decomposition of wood or other vegetable material without free access of air. The process has been going on for ages. It will be more fully treated in the succeeding volumes of this series. Ordinary coals are classed as *hard* and *soft*, or *anthracite* and *bituminous*. Then there are substances in which the decomposition has not progressed so far, such as *lignite* and *peat*.

I must call your attention again, and particularly, to the fact that charcoal unites with oxygen without

flame, and while it is in the solid state. The solid particles of the carbon as they burn with the oxygen become glowing white or *incandescent*, and thus give rise to the brightness of all ordinary flame. If it were not for the presence of incandescent or white-hot particles of carbon, our hearth-fires, and lamps, and candles, and gas-jets would all be reduced to a feeble sort of flame, such as that given by burning alcohol. This follows from a very important principle, namely, that whenever a substance burns—as iron burns in oxygen, and as carbon burns in Edison's electric lamp—without first assuming the vaporous state, it becomes exceedingly luminous. In fact, when any substance is intensely heated it becomes intensely bright, whether it burns or not, if it retains its solid state. If a candle or lamp does not burn well it gives off a quantity of smoke. This is merely unburned carbon. It is this very smoke which gives brilliance to the light if it does burn well. In that case the particles of smoke, which are carbon, become incandescent, and give a beautiful light, while at the same time they combine with oxygen in the form of an invisible gas and pass away into the air. You may perform a simple but instructive experiment to illustrate the part that carbon plays in giving brilliancy to flame.

To show you how abundant carbon is, and how it exists in places where you would perhaps least expect to find it, we will procure the carbon for this experiment from a lump of sugar. Lay a piece of writing-paper on a fire-shovel, place a lump of sugar on it, and hold

it over glowing coals. Sugar is nothing but carbon and water. You will presently see the sugar melt; then the water will pass off in the form of vapor—which you can collect and examine—and finally you have left a light, porous, cindery, black mass, which is carbon in the form of charcoal. Thus far we have merely driven off the water from the sugar by means of heat.

We might have separated it equally well by mixing sugar in the form of syrup with sulphuric acid. The acid would have combined with the water, leaving the carbon in a black mass.

Now fasten a little splinter of wood by a fine wire to your mass of charcoal, set fire to the wood so as to start the combustion, and then plunge the charcoal into a jar of oxygen. You will observe the brilliant effect of the combustion, without flame, owing to the bright incandescence of the particles of carbon as they combine with the oxygen.

The dazzling electric light, both incandescent and arc, is produced by similar glowing particles of carbon. In the Edison lamp, a slender thread, or filament, of carbon is raised to a white heat by the electric current; in the arc-lamp, the electricity heats the particles at the ends of the carbon-pencils, and the white hot sparks of carbon flying over from one pole to the other and uniting with oxygen as they go, produce the intense light.

Thus the blackest thing in the world becomes the source of our whitest light.

After learning this, you will perhaps not be so greatly

astonished to learn that although this amorphous kind of carbon, known as charcoal or lamp-black, is so dull and black, the crystallized form of carbon is the most splendid mineral in the world—the flashing diamond.

The lustre of the diamond is so intense that a special adjective is used to describe it, "*adamantine.*" The transparency of the diamond is perfect, unless, as often happens, there is some impurity in the stone. A very slight admixture of a foreign substance is sufficient, however, to cloud its clearness, dim its splendor, and change its color. Owing to such impurities diamonds show various tints of color, such as gray, yellow, green, and brown; and less frequently orange, red, blue, or even black.

The surest test of a diamond, next to that of its chemical properties, is its hardness. It is the hardest known substance; and can be scratched, polished, and cut only by its own powder. Its specific gravity is about 3.6.

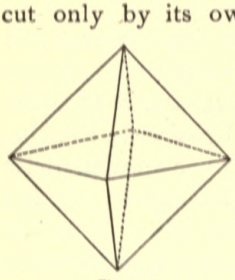


Fig. 19.

The usual form of the diamond crystal is a solid, bounded by eight equal and equilateral triangular faces (Fig. 19) called an *octahedron.*

The art of cutting diamonds was not known in Europe until after the middle of the fifteenth century, when it was invented by Louis Van Berguen, of Bruges.

India has long been celebrated as the home of the

finest diamonds, and in India the most famous locality for them has been Golconda; though, in fact, they are merely cut and polished there, being generally found in the districts lying further south and east. Diamonds are also found in Malacca, Borneo, and other parts of the East, in Brazil, Australia, South Africa, Russia, and the United States of America. Small and inferior

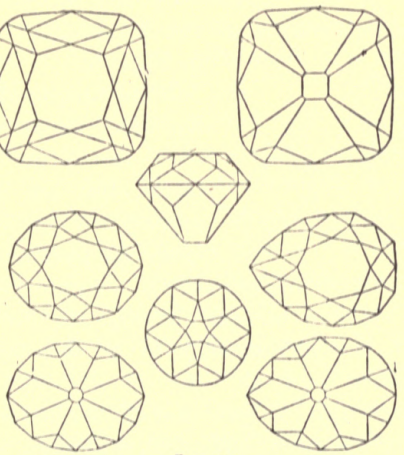


Fig. 20.

diamonds, and the fragments left when large stones are cut, are called "bort," and command a ready sale for use in the arts, being reduced to "dust" in a steel mortar, and used by lapidaries for cutting and polishing all sorts of gems. Minute fragments of bort are used for making the fine drills by which small holes are pierced in the jewels of watches. The use of small

diamonds for cutting glass is well known. For this purpose they are so mounted as to act upon the glass, not by a sharp corner, but by a rounded edge.

Diamonds are cut into various forms, but principally into *brilliants* and *rose diamonds*. The brilliant cut is the most expensive, difficult, and beautiful. It has an upper face or "bezel," which is octagonal, and this is surrounded by many facets—and as a rule the more the better. Some varieties of the brilliant cut are shown in Fig. 20.

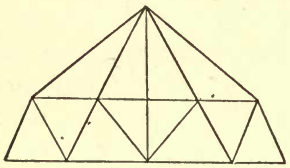


Fig. 21.

Rose diamonds have a flat base, above which are two rows of triangular facets, the six uppermost uniting in a point to form a sparkling pyramid. Fig. 21.

The largest diamond known weighs 367 carats—more than 2½ ounces—and belongs to the rajah of Mattan. It is egg-shaped, with an indented hollow near the smaller end.

Many years ago the governor of Borneo offered for it two war-brigs fully equipped, a number of cannon, a quantity of powder and shot, and $500,000 in cash; but the rajah refused to part with it, owing to a superstition which connects it with the good fortune of his family and the health of his province.

The Koh-i-noor, once a boasted possession of the great Mogul, now belongs to the English crown. It is said to have weighed 900 carats in the rough, but now has been reduced to 123.

The Regent, or Pitt diamond, weighing as cut 136¾ carats, is unrivaled in form and clearness. It was found in Golconda, brought from India by the grandfather of Pitt, the first Earl of Chatham, and by him sold to the duke of Orleans for £13,000. It decorated the hilt of the first Napoleon's sword of state, was captured by the Prussians at the battle of Waterloo, and now belongs to the king of Prussia. The Vanci diamond, weighing 106 carats, belonged to Charles the Bold, duke of Burgundy, who wore it in his hat at the battle of Nancy, where he fell.

A Swiss soldier found it, and sold it to a clergyman for a gulden—about half a dollar. It passed into the possession of Anton, king of Portugal, who sold it for 100,000 francs ($20,000). Later it became the property of a French gentleman named Vanci.

A descendant of Vanci being sent as ambassador was required by king Henry III. to leave the diamond as a pledge. A servant was accordingly sent to carry it to the king, but on the way he was attacked and murdered by robbers, not, however, before he had contrived to swallow the diamond unobserved. His master, confident of his fidelity, caused the body to be opened, and recovered the jewel. This diamond came into the possession of the Crown of England, and James II. carried it to France in 1688. Louis XV. wore it at his coronation. In 1835 it was purchased by a Russian nobleman for nearly half a million dollars.

The following interesting account of the South African diamond mines is condensed from the report of George F. Kunz, for 1887.

"The author of the Arabian Nights undoubtedly thought he was imagining the most improbable thing when he described the 'Valley of Diamonds,' found by Sindbad the Sailor. Yet when compared with the African mines that valley pales into insignificance.

"The primitive method of washing was carried on for centuries by thousands of slaves driven by the merciless whip of their master; but improved methods have gradually been introduced. Steam railroads were run into the mine. Millions of tons of reef were removed. Dynamite was substituted for powder.

"After the broken earth is raised, it is put on the 'sorting-ground,' where it is partly disintegrated by water and the weather. After being more finely broken up, it is passed down into large vats containing immense wheels, by which the rock is finely divided. So accurately and powerfully does this machinery work, that all the diamonds, even those of the size of a pin-head, are saved. Thirteen million gallons of water are annually hoisted from the Kimberly mine. Ten thousand natives, 2,500 horses, mules, and oxen, and 350 steam-engines are employed.

"During the last ten years, the South African mines have yielded 27,878,587 carats, valued at £31,717,341.

"Before the organization of the great companies which now work these mines as systematically and economically as a woolen-mill, it is believed that about one-quarter of the diamonds found were stolen by the workmen. The natives use the most ingenious methods for concealing the precious gems. On one occasion

some officers, suspecting that a Kafir had stolen diamonds, gave chase and caught him, just after he had shot one of his oxen. No diamonds were found upon him, but he got off with them nevertheless, for he had cunningly loaded his gun with them, and after the departure of the officers he had the satisfaction of digging them out of the body of his dead ox.

"More diamonds weighing over 75 carats after cutting have been found since the African mines were opened than were ever known before. On March 28, 1888, there was found in the De Beers mine an octahedral crystal of diamond weighing 428½ carats.

"It was valued at $15,000. From its form it is believed that it can be cut into a brilliant of 200 carats; hence it will be the largest known."

Besides the forms already described, namely, charcoal, coke, lamp-black, bone-black, coal, lignite and diamond, carbon also occurs in the form known as graphite, plumbago, or "black-lead," which is used by every one in ordinary pencils.

Of course there is no lead about it; it is only a form of carbon. It can be artificially prepared by dissolving charcoal in melted iron, from which on cooling the carbon is deposited as graphite. Graphite has a grayish black streak, and a metallic lustre, although it is not a metal. This would be indicated by its brittleness, as well as by its infusibility.

It has a greasy or unctuous "feel," and is largely used as a lubricant. It is not found pure, but contains a small admixture of iron, silica, lime, alumina, etc. It

sometimes occurs in tabular or scaly hexagonal crystals, but usually massive. It is much less combustible than hard coal, and burns with difficulty, even before the blow-pipe, so that it is very useful for the manufacture of melting-pots or "crucibles," which have to withstand great heat. For this purpose, however, it is mixed with half its weight of clay. Graphite is the chief ingredient in stove-polish.

QUESTIONS ON CHAPTER 10.

1. In what substances is carbon found?
2. What are the three leading forms of carbon?
3. Describe charcoal and its uses.
4. Mention other similar forms of carbon.
5. Explain the brilliancy of ordinary flame.
6. Of what is sugar composed?
7. Explain the brightness of the electric light.
8. Describe the diamond.
9. What is the usual form of its crystal?
10. What is "bort?"
11. Describe graphite. Its uses?
12. Form of its crystals.

CHAPTER XI.

CARBON, CONTINUED.

In the preceding chapter we studied carbon, one of the most abundant, common, and wonderful elements. We found it appearing, as if by magic, under various forms of the utmost dissimilarity: as soft as velvet and as black as night in lamp-black; harder than steel, and as bright as day in the diamond; now the chief constituent in our illuminating oils and in our fuels, on account of its brilliant and fiery combustion with oxygen; and again used for crucibles, because in the form of graphite it will neither burn nor melt; in the form of "bort" we saw it used to cut and scratch, and rasp the hardest substances, and again as graphite we found it used instead of oil to prevent the friction of wheels and shafts. It is the hardest and the softest, the dullest and the brightest, the cheapest and the dearest, the roughest and the smoothest; in a word, the most contradictory element in nature.

How then do we know its identity? Why are we certain that charcoal and graphite and the diamond are not really three distinct minerals instead of three forms of the same mineral?

In the first place, notwithstanding these marked differences there are also certain marked resemblances be-

tween these forms of carbon. They are insoluble in all known liquids; they are tasteless, inodorous, and infusible; and when heated without free access of air, they remain unchanged unless the temperature is very high.

In the second place, we know that they are varieties of the same substance, because if we burn the same weight of diamond, graphite, and charcoal, we obtain in each case the same substance, "carbon dioxide," CO_2.

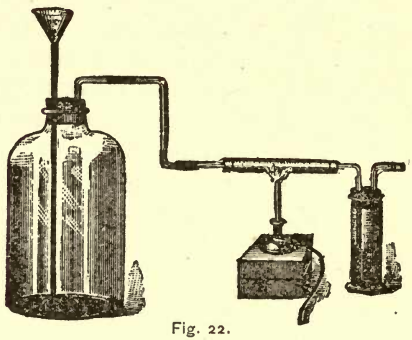


Fig. 22.

In each case the oxygen of the product comes from the air, and in each case the weight of the carbon in the product is exactly equal to the weight of the diamond or graphite or charcoal consumed.

I will not ask you to verify this statement by burning a diamond, unless you have plenty of them to spare; but that experiment has repeatedly been performed, and always with the same result. You may, however, with much profit make the experiment of combining char-

coal with oxygen, in the following manner: Put a small piece of charcoal into a hard glass tube. Heat this tube by an alcohol flame, or a gas-jet, at the same time passing oxygen through it. Pass the resulting gas into clear lime-water. Arrange the apparatus as shown in Fig. 22.

A is a bottle containing oxygen; B is a hard glass tube containing the charcoal; C is the jar of clear lime-water. The reaction which takes place is represented thus:

CaO_2H_2	+	CO_2	=	$CaCO_3$	+	H_2O
Lime-water.		Carbon dioxide.		Calcium carbonate.		Water.

That is, the oxygen passing over the heated charcoal combines with it to form a new gas, and this shows itself on passing into the lime-water by combining with the lime in solution, forming carbonate of lime, which is insoluble in the water, and therefore sinks as a white chalky powder. Now, no other gas except carbon dioxide acts in this way with lime-water.

Hence we may conclude that the gas formed is carbon dioxide. This is a useful test, and whenever under ordinary circumstances an unknown gas passing into lime-water produces an insoluble substance, we may conclude that the gas is carbon dioxide.

Notice in passing the significance of the chemical termination—*ate*. I have already shown you that the termination—*ide*, is used in names of the compounds of *two* substances. Thus every *oxide* is a compound of oxygen with one other element; every *sulphide* is a

compound of sulphur with one other element, etc. Now the termination—*ate*, is used in names of the compounds of *three* substances, one of which is always oxygen, and another of which is always indicated by that part of the name which precedes the termination—*ate*. Thus every *carbonate* is a compound of oxygen, carbon, and one other substance. Every *sulphate* is a compound of oxygen, sulphur, and one other substance, etc. The name of the third element in every such compound is written after, and connected by the word "of;" thus, carbonate of calcium; sulphate of lead, etc.; or it may equally well be written *before*, as calcium carbonate, lead sulphate, etc. What is the composition of "silver nitrate?"

Now, by the lime-water test, as I was saying, it has been proved that charcoal, graphite, and the diamond are all varieties of the same substance, carbon, because, all of them when combined with oxygen form a gas which when passed into lime-water combines with the lime and forms the insoluble substance, carbonate of lime. This property of carbon, and other elements, by virtue of which they exist in dissimilar forms, is called *allotropism*.

Before leaving our study of carbon we must notice that its affinity for oxygen is so great that, under proper conditions, it is able to abstract the oxygen from oxides.

This you may prove by the following experiment:

Mix two or three grains of powdered copper oxide, CuO, with about one-tenth its weight of powdered

charcoal. Heat the mixture in a tube fitted with a small outlet tube as shown in Fig. 23.

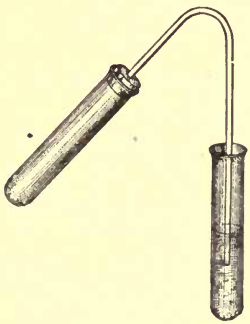


Fig. 23.

Pass the gas which is given off into clear lime-water, contained in a test-tube. Is it carbon dioxide? What evidence have you that oxygen has been removed from the copper oxide?

Does the substance left in the tube suggest the metal copper?

The reaction between the charcoal and the copper oxide is represented thus: $2CuO+C=2Cu+CO_2$.

This abstraction of oxygen from a compound is called *reduction*. Hence carbon, which effects it, is called a *reducing agent*. It is extensively used in extracting metals from their ores, which are usually oxides. Thus iron occurs in nature, not pure, but usually in combination with oxygen. In order to get the pure metal the oxygen must be extracted. This is accomplished by heating the ore with some form of carbon, either charcoal or coke.

To conclude this lesson, let us study that most important compound of carbon and oxygen, which we have already discovered and learned to recognize as carbon dioxide, CO_2. It is commonly called *carbonic acid*. It has already been shown that carbon dioxide is formed by the combustion of charcoal. In a similar

manner it can be shown that this gas is formed whenever any ordinary material is burned. For example, hold an empty bottle over the chimney of a burning lamp so that it may receive the products of the combustion of the oil.

You will soon find something besides ordinary air in the bottle. Pour a little clear lime-water into it, and shake it up. Do you see any evidence of the presence of carbon dioxide?

Burn a piece of paper in a jar, pour in a little lime-water, and observe the result.

Consider now what is constantly going on in your own body when you breathe. Perhaps you have never thought of it at all; perhaps you have supposed that in breathing you simply draw a quantity of air into the lungs and breathe the same quantity of air out again. Try this experiment: Arrange a bottle or flask of lime-water as in Fig. 24.

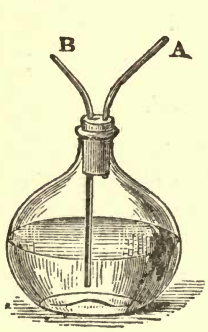


Fig. 24.

Placing your lips on the tube *a*, you can either draw the outside air through the lime-water, or blow the air from your lungs through it. Try first the experiment of drawing the outside air through the water. Observe that so long as you do this no effect is produced upon the water. Now throw the air from your lungs through the lime-water several times in succession. See how the lime-water grows chalky. This teaches you a very important lesson.

The air that you breathe out is no longer pure, but contains *carbon dioxide.* A sort of combustion is going on in your lungs; the oxygen of the air is combining with a portion of the carbon contained in your blood, forming carbon dioxide, which is expelled from your body when you breathe. This combustion of oxygen and carbon in the lungs is one of the means by which the body is so wonderfully heated, that in summer and winter it is maintained at the most uniform temperature, and with the least expenditure of fuel. In twenty-four hours the lungs of a man convert seven ounces of carbon into carbon dioxide.

Just here, I must call your attention to another property of this gas. Into a jar of carbon dioxide plunge a lighted taper. It is at once extinguished. If a mouse or bird be dropped in the gas, its life goes out like the flame of the taper. One or two breaths from your lungs will furnish enough gas for this experiment.

Set a bottomless jar or bottle in a vessel of water. Close it with a tight cork, in which a small tube is closely fitted.

Now inhale the air from the jar, and breathe it back, as shown in the figure. Even one inspiration is enough to spoil the air so that a taper will not burn in it. This forcibly illustrates the necessity of an abundant supply of fresh air in our homes and schools and churches. It proves the need and indicates the character of good *ventilation.*

The carbon dioxide must be removed as fast as it is breathed out, and plenty of air containing its full measure of oxygen must be supplied.

It must not be supposed, however, that carbon dioxide is a poisonous gas any more than nitrogen is.

Animals die in it merely because of a lack of oxygen; they suffocate, as they do in water or as they would in nitrogen. Most of the bad effects of breathing the air of ill-ventilated rooms, such as headache, drowsiness, etc., are caused by other impurities contained in the breath.

The total amount of carbon dioxide formed by breathing is almost alarming. Faraday estimates that the lungs of London alone pour out 548 tons of this gas every twenty-four hours!

Add to this the vast quantities that result from all the combustion that is constantly going on, and you will wonder that in the course of ages the whole atmosphere does not become unfit to breathe. What prevents it?

Just here comes in one of the most beautiful and wonderful contrivances of nature. This gas so deadly to animal life is the very life and support of plants. All the plants that grow upon the earth absorb carbon from the air. By the aid of sunlight and sun heat, they decompose the carbon dioxide which animals breathe out, and build up from it the complex compounds of carbon which form their tissues, and set free again the oxygen that is so necessary for animal life. Give to plants a pure air like that which is best for us, and they cannot live in it; give them carbon dioxide and other matters and they live and rejoice.

From this our dependence upon the sun is made evi-

dent. Every living thing is dependent upon the decomposition of carbon dioxide by plants. This decomposition cannot be effected without the aid of the sun. If the sun should stop shining, all earthly life would cease.

We have seen that when carbon dioxide passes into lime-water it combines with the lime to form carbonate of calcium, or chalk. As it is a poor rule that will not work both ways, we ought, therefore, to be able to decompose chalk, or any form of carbonate of lime, and get the carbon dioxide again. And this we can easily do. If some moistened chalk be put into a retort and heated red-hot carbon dioxide is driven off.

A simpler method of decomposing carbonates, however, is by the use of an acid.

Into a jar containing a few small chippings of chalk, limestone, marble, or oyster-shell—all varieties of carbonate of calcium—pour a little hydrochloric acid. A brisk boiling—in this case called *effervescence*—immediately begins.

It is not steam, however, that is bubbling up, but a gas. Test the contents of the jar with a taper. The flame is instantly extinguished. Pass some of the gas into lime-water; the chalky carbonate is formed.

Here then we have again carbon-dioxide, the very same gas we formerly produced by the direct combination of carbon and oxygen; and by this action of the acid upon a carbonate we obtain the gas more readily, and in great abundance.

Having it, now, in sufficient quantity, familiarize

yourself with it by as many experiments and tests as you can invent.

Ascertain whether it is heavier or lighter than air. See whether you can pour it out of a jar as you can water. Pour some on a candle-flame. Balance a jar on a pair of scales and pour some of the gas into it. Does it affect the scales?

Carbon dioxide dissolves in water, one volume of gas dissolving in about its own volume of water at the ordinary temperature. When the pressure is increased more gas dissolves; and when the pressure is removed the gas again escapes. The "soda-water" of the druggists is simply water charged with carbon dioxide under pressure.

Much of the life and sparkle of ordinary spring water is due to carbon dioxide dissolved in it. Spring water is therefore a sort of natural soda-water. Rain-water, like boiled or distilled water, has a flat, insipid taste, but as it trickles down the mountain-side, and dashes from ledge to ledge of rock, it dissolves the gases in the air and becomes thus naturally aerated.

QUESTIONS ON CHAPTER II.

1. How can the identity of charcoal, graphite, and diamond be proved?
2. What is the product of the combustion of carbon with oxygen?
3. What effect has this gas on lime-water?
4. What is the significance of the termination "*ate?*"
5. What is *allotropism?*

6. What can you say of carbon as a "reducing-agent?"

7. How is air changed by breathing?

8. Describe carbon dioxide.

9. What is the cause of the necessity for ventilation?

10. Why does not the whole atmosphere become tainted with carbonic acid?

11. How can carbon dioxide be obtained from limestone or chalk?

12. What is "soda-water?"

CHAPTER XII.

A PIECE OF MARBLE.

You will remember that we were led to our study of carbon by finding bits of charcoal in a handful of earth which we heated. In this earth we have also found water and quartz.

Let us now examine it further: Dry a tablespoonful of garden-soil without burning it, and pour a few drops of hydrochloric acid upon it. If your garden is like mine you will notice a brisk effervescence; which will remind you of the experiment described in the preceding chapter, by which we obtained carbon dioxide by treating carbonate of lime with hydrochloric acid.

We then used lime dissolved in water as a test for the carbon dioxide, and decided that any gas which passing into lime-water at ordinary temperature renders it chalky is carbon-dioxide; and we learned that the chalk thus formed is carbonate of calcium.

With this knowledge we may now work backward, and make the production of carbonic oxide by the action of hydrochloric acid upon a mineral a test for carbonate of calcium; for this is the only common mineral which effervesces freely with dilute hydrochloric acid at ordinary temperatures.

We, therefore, may reasonably conclude that our

soil contains calcium carbonate, and proceed to study this substance in some detail. It will be convenient, as in the case of quartz, to use larger specimens than the minute grains mingled with the soil, and we will therefore take a piece of white marble—specimen No. 20 in the collection.

Repeating the experiment with hydrochloric acid we obtain again a plentiful supply of carbon dioxide, which we allow to escape, however, as we are now more interested in the other elements in the carbonate. The reaction is expressed as follows:

$$CaCo_3 + 2HCl = CO_2 + H_2O + CaCl_2.$$

This means that calcium carbonate and hydrochloric acid give, besides the carbon dioxide, water and calcium chloride. The chloride is dissolved in the water.

Evaporate the solution to dryness.

The colorless needle-shaped crystals remaining are calcium chloride, still combined with a portion of water. In this condition it is called *hydrated* calcium chloride. Expose this substance to the air. Observe how rapidly it becomes moist by absorbing water from the air. This property makes it valuable for drying gases. In many experiments, gases containing undesirable moisture are passed through tubes containing calcium chloride, which absorbs the water and leaves the gases dry. If hydrated calcium chloride be strongly heated it fuses and parts with all its water, becoming simply $CaCl_2$.

Calcium chloride consists of two very important ele-

ments, *calcium*, which we are now specially in search of, and *chlorine*, a gas which we shall consider later.

Unfortunately it is not easy to separate them. I say "unfortunately," for it is important for you to become familiar with the leading elements in their purity, as far as possible. The analysis of calcium chloride, though difficult, is not impossible, and is effected by fusing it, and passing a powerful current from an electric battery through it. If you have some friend who is an expert chemist or mineralogist, it will be worth your while to get him to perform the experiment for you. You will then see the calcium separating from the chlorine in the form of minute drops or globules. It is a yellowish-white metal, between gold and lead in hardness. It gradually grows dull, or *tarnishes*, on exposure to the air.

In order to keep it bright it is kept under naptha. It can be rolled into sheets and hammered into leaves. At a red heat it melts and burns with a dazzling light, accompanied with showers of sparks. When brought into contact with water it unites strongly with the oxygen, forming lime, CaO, and setting the hydrogen free. This is another proof of the composition of water, and is a very interesting and instructive experiment. You remember that we have already seen that the oxygen of water will desert its comrade, hydrogen, for the sake of uniting with red-hot iron. The attraction between oxygen and calcium is so much greater that oxygen combines with it at ordinary temperature.

Returning now to our marble, $CaCO_3$, we will try

to decompose it without the help of acid. Put a piece of marble into a stove or furnace, and leave it over night. On taking it out you at once perceive a decided change in its appearance, and a more decided change in its properties. It has, in fact, been converted into *quicklime*, the carbon dioxide having been driven off by heat.

The decomposition is represented thus:

$$\underset{\text{Calcium Carbonate.}}{CaCO_3} = \underset{\text{Quicklime.}}{CaO} + \underset{\text{Carbon dioxide.}}{CO_2}$$

Expose, now, one-half of the quicklime to the air, and place the remainder in a cup half full of water.

The portion exposed to the air gradually absorbs both moisture, H_2O, and carbon dioxide, CO_2, and is resolved again into calcium carbonate, in the form of a white powder. This is commonly called *air-slaked* lime.

The portion put into the cup rapidly combines with the water, becomes very hot, crumbles into a fine powder, and is converted into the *hydroxide* or *hydrate* of calcium. The result of this *water-slaking* is thus represented:

$$CaO + H_2O = CaO_2H_2$$

This substance is partially soluble in water, and forms *lime-water*. Pour the slaked lime from your cup into a jar, or wide-mouthed bottle, containing perhaps a pint of water. The undissolved lime will settle to the bottom. After some hours the solution above will become clear. Pour this off carefully for experiment.

What takes place when some of the solution is exposed to the air? When the gases from the lungs are passed through it? When carbon dioxide is passed through it? When dilute sulphuric acid is added?

Slaked lime mixed with sand is used for mortar, whitewash, and plastering. When the mortar is spread and exposed to the air it takes up carbon dioxide, and thus the calcium oxide, or lime, becomes calcium carbonate again, and hardens into stone.

Notice right here a curious property of carbon dioxide. We have repeatedly seen that a small quantity of it, passed into lime-water, unites with the lime, forming the chalky carbonate of calcium. Now repeat that experiment, but allow the gas to pass into the lime-water for a longer time. You will be surprised to observe that the chalky water gradually grows clear again. The explanation of this phenomenon is found in the fact that water saturated with carbon dioxide dissolves the carbonate. To this property we owe the formation of caves in limestone regions with their beautiful stalactites which depend like icicles from the roof, and the curious stalagmites which stand up like inverted icicles from the floor. Water containing carbon dioxide trickles through crevices and dissolves away the limestone little by little until in the course of ages large caves are formed. Then the water holding calcium carbonate in solution drips slowly from the roof, and as the water evaporates little particles of the carbonate are deposited, until the long glistening stalactites are formed; while the drops that fall to the floor build up

in like manner the crystalline stalagmites in many fantastic shapes.

The following simple experiment reveals a property of lime or calcium oxide which you might not otherwise observe; namely, its power in counteracting the effect of an acid.

The most common test for acids depends upon the fact that they instantly change the color of certain substances. One of the substances whose color is most easily and noticeably changed by an acid is called *litmus*. Litmus is a coloring matter which is obtained from several sorts of lichen or fungus. The lichens are powdered and steeped in ammonia-water until they are decomposed.

After further treatment, not necessarily described here, a peculiar purple dye is obtained, which is the *litmus* of commerce.

The interesting fact about this litmus is that upon the addition of any acid it turns *red*; and again, upon the addition of lime or any alkali, it turns *blue*. This you should now prove by several experiments. You will find in the collection accompanying this book (specimen No. 21) a piece of blue "litmus-paper," which is simply a piece of ordinary unsized paper, dyed with an alcoholic solution of litmus and treated with some alkali.

You can turn a piece of this blue paper red by moistening it with an acid, or more easily by moistening it and holding it over burning sulphur; a burning match will answer the purpose. You can get this paper, both red and blue, at a drug-store, or, better, you can buy litmus and prepare it yourself.

Now, to see the effect of lime, moisten a piece of red litmus-paper, and place a few grains of quicklime upon it; it turns blue. Try the effect of a little powdered marble.

Now heat some of the marble red-hot, either by laying it on a piece of charcoal and applying a blow-pipe flame, or by holding it in the flame with a pair of forceps.

Place this *burned* marble upon the moistened red paper, and note the effect. Here is another proof that calcium carbonate, when heated, parts with its carbon dioxide, and become calcium oxide, or quicklime.

Instead of using the paper, you may try the same experiments with more striking results upon a glass full of the red solution of litmus.

It is to this *antacid* action of the lime that it owes its valuable medicinal properties, lime-water being one of the best remedies for acidity and irritability of the stomach.

In the ashes which remain after the combustion of any vegetable matter a proportion of lime is found. This shows that calcium oxide is one of the most important plant-foods, and indicates the use of lime as a fertilizer.

The carbonate is sometimes applied without previous preparation, in the form of marl or chalk; but usually after having been calcined, and reduced to the powdered oxide by slaking.

The quantity of calcined lime used varies from three to eight tons to the acre; the larger quantity being re-

quired for "strong" land, or land holding much vegetable matter. The fertilizing effect of lime is due even more to its chemical effects upon the soil than to the food it directly affords to the crops. It promotes the decomposition of all kinds of vegetable matter in the soil, and corrects any acidity present, and thus tends to destroy those weeds which flourish in an acid soil. On certain kinds of soil, the finer grasses do not thrive until the land has been limed. Lime is the only reliable cure for "finger-and-toe," "club-root," "ambury," in turnips.

In the preceding chapter we learned that the brilliancy of the electric light is due to the incandescence of some highly heated infusible substance, usually carbon.

Lime is infusible, and when strongly heated yields a light fairly rivalling the electric lamp. A jet of oxygen and hydrogen burning together in an intensely hot flame is directed against a cylinder of lime. The light produced in this way has been seen for more than one hundred miles. It is called the "Drummond light," from its inventor, Captain Thomas Drummond.

To calcium carbonate, the mineralogists have given the name *calcite*. In nature calcite occurs as marble, chalk, limestone, shells, corals and pearls, and in a crystallized form known as calc-spar.

In all its forms it is readily distinguished from quartz by its inferior hardness, and by its effervescence with acids. Its hardness is 2.7. Shells, corals, and pearls, being evidently animal products, will more properly be considered in the succeeding volume of this series.

Common limestone, chalk, and marble are also largely derived from organic structures, but for the present they may be considered as ordinary minerals.

Limestones are either compact or granular, the compact form breaking with a smooth surface, often shell-like or conchoidal. Granular limestones, of which, indeed, marble is a fine variety, are crystalline, and break with a more or less sugary fracture. Ordinary limestone is one of our most important building materials.

Marble is capable of a high polish, and is familiar to all in the form of table-tops, mantels, floors, and monuments.

The purest and finest marble is used in sculpture, and is called statuary marble. The finest qualities come from Carrara, in Italy, and from the Island of Paros, whence the celebrated Parian marble of the Greek sculptors takes its name.

Many ancient temples were built of marble, notably the Parthenon at Athens. The capitol at Washington is built of white marble, much of which came from the quarries at Lee, Massachusetts. Excellent marble is found abundantly in Vermont, Western Massachusetts, Eastern New York, and Western Connecticut. Berkshire County, Mass., rests on a marble and limestone floor, at least a thousand feet in thickness.

Many varieties of marble are beautifully shaded, veined, and mottled.

Chalk is a soft earthy variety of limestone, forming great layers, or "strata," in the earth, and from its animal

origin and structure proving of even more interest to the geologist than to the mineralogist. It is generally of a yellowish tinge, but sometimes of snowy whiteness. It is easily broken, has an earthy fracture, is rough to the touch, and clings slightly to the tongue. It generally contains a little silica, alumina, or magnesia. It is sometimes so compact as to be available for building.

It can be burned into quicklime, and nearly all the houses in London are cemented with mortar made of chalk.

When freed from its particles of silica by pounding and washing, it becomes "whiting," familiar to every housewife. Carpenters and others use it for making marks which are easily erased. The blackboard with its chalk or crayon has come to be as familiar in the lecture-room of the university as in the humblest village school.

Calc-spar is the crystallized variety of calcite, and it occurs in several hundred shapes, all modifications of the rhombohedron, which, you remember, is also the regular form of an ice-crystal (specimen No. 22). This represents a variety very abundant at Lockport, N. Y., and known as "dog-tooth spar." All varieties of calc-spar, split or "cleave" into rhombohedrons. Transparent specimens have the property of making objects viewed through them appear double. This property is called *double refraction.*

QUESTIONS ON CHAPTER 12.

1. What is the effect of hydrochloric acid upon carbonate of lime?

2. What is the chemical composition of marble?

3. What other substances have the same composition?

4. How can the element calcium be obtained? Describe it.

5. What effect has its metal upon water?

6. How can marble be decomposed without acid?

7. What are the properties and uses of lime?

8. Explain the Drummond light.

9. What is litmus-paper?

10. What is calcite? Name several varieties.

11. Why is the carbonate of calcium of interest to the geologist?

12. What is chalk?

13. Describe crystallized calcite.

CHAPTER XIII.

CLAY.

The most noticeable substance remaining in our handful of earth, now that we have removed the water, quartz, charcoal, and limestone, is the clay that we found binding these substances together.

Clay owes its origin to the decomposition of minerals and rocks, and is rather a mixture of the fine particles of other disintegrated substances than a simple mineral. We shall, therefore, reserve our study of its formation until we come to treat of geology, but it is of present interest to us on account of the decomposed feldspar which enters largely into its composition, and particularly on account of the metal, *aluminum*, which next to silicon is the leading element in this feldspar. Clay is essentially a silicate of aluminum. Clay consisting of pure aluminum silicate is called *kaolin*.

Clay is a kind of soil marked by smoothness and stickiness, or "tenacity," and by the fact that it can be "worked" or molded into any desired shape. It is used by sculptors in modeling the forms which are afterward to be cut from marble or cast in bronze. On baking it becomes as hard as stone, and is therefore of great value for the manufacture of brick and tiles.

The various forms of porcelain and earthenware are

made of baked clay, covered with some substance which melts at a high temperature, and forms a glaze or enamel binding the clay together, giving it a smooth surface, and rendering it impervious to liquids and air, by filling up the pores which abound in clay baked without an enamel.

For the manufacture of porcelain the finest white or china clay is used, while for common earthenware colored clays may be used. The glaze used for porcelain is usually finely powdered feldspar. For ordinary earthenware, a "salt-glaze" is used. This is made by throwing common salt into the furnaces in which the ware is baking. The salt, decomposed by heat, causes a deposit of fusible silicate upon the surface of the clay.

When clay becomes firmly compacted it becomes *slate*.

The peculiar clayey, or *argillaceous* odor of slate is enough to indicate its identity with clay, even if we did not know the history of its formation. We find many degrees of hardness in clay. It sometimes becomes very hard by simple drying, as we often see in dry weather, but this is not slate, and no amount of mere drying will convert clay into slate, for dried clay when moistened with water is easily brought back to its putty-like or *plastic* state. To make a good slate the hardening must be the result of pressure, as well as dryness, aided probably to some extent by heat. True slate is a permanently hardened, but not burnt clay, which will not become soft when wet.

Slate is easily scratched with a knife, and is readily distinguished from limestone, because it does not effer-

vesce with acid. It naturally shows the same varieties in color and composition as clay. A good assortment of colors is afforded by the roofing-slates. Specimen No. 23 is a typical slate, for it not only has a compact structure, and an argillaceous odor, but it is very plainly formed in layers, or, as geologists say, it is distinctly *stratified.* This means not simply that it splits into leaves, but that it was formed by the deposition of layers of clay, held in suspension in water. Ordinary roofing-slates, like specimen No. 24, rarely show true stratification. The thin layers into which they split have been developed by pressure after the formation of the slate by deposition, and they do not follow the lines of the real stratification. This structure is known as "slaty cleavage." Some roofing slates, known as "ribbon-slates," show bands of various colors across the flat surfaces. These bands indicate the lines of the true bedding, or natural layers of deposition.

Slate which splits easily into thin, brittle layers, *parallel with the bedding*, is known as *shale.*

We must now consider the mineral feldspar, to which the origin of most of the slates and clays can be traced. It must be borne in mind, however, that *feldspar* is the name of a group of minerals, rather than of any one species. Geologically, the feldspars are the most important of all minerals for more than any others, with the possible exception of quartz, they enter into the composition of *rocks*, which, by the way, are large mixtures or aggregations of minerals, no single crystal or mineral-grain being properly called a rock.

The abundance of the feldspars is well expressed by the name—feldspar being simply the German for field-spar, and implying that it is the common spar or mineral of the fields.

Like clay, the feldspars are essentially silicates of aluminum, with the addition of smaller quantities of other elements.

Their general characteristics, including easy cleavage in two directions at right angles to each other or nearly so, are well exhibited in the common variety, *orthoclase*, specimen No. 25.

This is the most abundant of all minerals, forming the principal part of granite, gneiss, and many other important rocks. Its most common colors are white, gray, pink, and flesh-red.

You may determine for yourself the physical properties of orthoclase, and we will postpone to a later chapter the examination of other varieties of feldspar.

Orthoclase is a silicate of aluminum and potassium; that is, it is composed of silicon, oxygen, aluminum, and potassium, its chemical formula being $K_2Al_2Si_6O_{16}$.

The metal aluminum cannot readily be obtained directly from feldspar or clay, and the various processes of its manufacture are too complicated to be explained in detail. In general, an aluminum sulphate is first formed by decomposing clay with sulphuric acid. This sulphate of aluminum is then combined with an alkaline sulphate to produce *alum*. Then, to a solution of the alum, *ammonia* is added, and a white powder settles

to the bottom of the solution. This powder on being heated becomes pure *alumina*, which is a compound of aluminum and oxygen, Al_2O_3, and is the only oxide of aluminum known. Finally, this *alumina* is mixed with carbon, and subjected to the action of a powerful current of electricity, which separates the metal aluminum from the oxygen. The apparatus used for this purpose is known as the Cowles electric furnace. There are other processes for obtaining this metal, but this is one of the best and cheapest.

The cheap production of aluminum from clay is one of the important practical problems needing to be solved, because the properties of this metal are such as to render it of great value, while its abundance in its natural compounds is so unlimited that nothing but the difficulty and expense of preparing it prevents it from taking (for many purposes) the place of iron, to which in some respects it is greatly superior. In the first place it is but little more than one-third as heavy, its specific being 2.7, while that of iron is 7.8; moreover, it has a beautiful color and lustre resembling silver, and it does not easily rust or tarnish on exposure to dry or moist air. It is exceedingly strong, and yet is so malleable that it can be hammered into thin sheets and wrought into any desired form.

A slight alloy of copper has been found to add yet more to its toughness and strength. Serious drawbacks to its usefulness, however, are its liability to be attacked by salt, and by vegetable acids, the difficulty of soldering it, and the difficulty of working its hard alloys.

Alumina, the oxide from which, as we have seen, the metal aluminum is obtained, occurs crystallized in nature in the forms of *corundum*, *ruby*, and *sapphire.*

Corundum is inferior only to the diamond in hardness. It is generally of a dull and muddy appearance, even in its hexagonal crystals.

A common variety is *emery*, which is universally known from its use in polishing and cutting gems, metals, and glass.

Ruby and sapphire are identical except in color, and are, indeed, varieties of corundum. No gems, with the possible exception of the diamond, are so highly prized. They are transparent and brilliant. The finest rubies, which are as red as fire, are found in Burmah and Siam, and the finest sapphires, of an exquisite blue color, come from Ceylon. In Burmah, when a very fine stone is found, a procession of officers and soldiers mounted on elephants is sent to receive it. One of the titles of the king of Burmah is "Lord of the Rubies." A ruby weighing more than twenty carats is commonly called a carbuncle. A perfectly pure crystal of corundum, transparent and colorless, is known as white sapphire, and has been mistaken for the diamond.

QUESTIONS ON CHAPTER 13.

1. What is clay? Its uses?
2. Describe slate, and its origin.
3. Distinguish between the "stratification" and "cleavage" of clay.

4. Describe orthoclase. Its composition?

5. How is aluminum obtained, and what are its properties?

6. Mention and describe three varieties of crystallized alumina.

CHAPTER XIV.

POTASSIUM—MICA.

In our study of clay we learned that it is essentially a hydrous silicate of aluminum; but besides silicon, oxygen, hydrogen, and aluminum, it sometimes contains another element, namely, *potassium.*

We might also have found traces of potassium in the ashes left mingled with charcoal by the burned sticks and roots, for when vegetable material is burned the potassium remains behind, chiefly in the form of potassium carbonate. When wood-ashes are treated with water, the potassium carbonate dissolves, and may be obtained, though in an impure state, by evaporating the solution. The substance thus obtained is popularly called *potash*, or *pearl-ash.* Its chemical formula is K_2CO_3.

Treat two or three pounds of wood-ashes with water. Filter the solution and test its effect upon red litmus-paper. You observe its strong alkaline reaction in turning the red paper blue.

Evaporate the solution to dryness, collect the residue, and heat it in a test-tube with a little hydrochloric acid. The gas given off you will recognize as carbon dioxide if you pass it into clear lime-water.

The metal potassium was first obtained by passing

a strong electric current through caustic potash, HKO. It is now prepared by heating potash and carbon together to a high temperature in an iron retort. The carbon and the oxygen of the potash combine and escape as a gas, hydrogen is set free, and the metal potassium, which vaporizes at red heat, distils over after the manner of water.

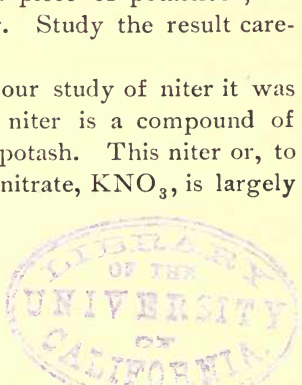
The preparation of this metal is attended with many difficulties and extreme danger, and should not be attempted except by experts.

When obtained in its pure state, potassium is a bright, silver-white metal, light enough to float on water, and so soft that it can be easily cut with a knife. It rapidly absorbs oxygen when exposed to the air, and gradually becomes converted into a white oxide.

It unites with the oxygen of water with great energy, setting the hydrogen free, and igniting it at the same time by the heat evolved by the combination. The flame of hydrogen thus ignited is tinged with a peculiar purple, which is characteristic of potassium.

In consequence of these properties of potassium it cannot be kept in the air, but must remain until desired for use under some oil, like petroleum, upon which it does not act. Throw a small piece of potassium, not larger than a pea, upon water. Study the result carefully.

You may recollect that in our study of niter it was incidentally mentioned that niter is a compound of nitrogen, with oxygen and potash. This niter or, to speak chemically, potassium nitrate, KNO_3, is largely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used, you remember, in the manufacture of gun-powder and fire-works.

One of our first experiments was the preparation of oxygen by heating together oxide of manganese and potassium chlorate. This potassium chlorate may be obtained by passing chlorine gas through a hot solution of caustic potash. It occurs in white rhomboidal crystals of a pearly lustre, has a cooling taste, and fuses easily. It parts so readily with its oxygen, that if substances with a strong affinity for that gas, such as carbon, sulphur, or phosphorus, are heated with it, even so slightly as by friction, they combine with the oxygen with explosive force. Chlorate of potash is therefore used in the manufacture of friction-matches, together with phosphorus and sulphur. We need not consider now any of the other numerous compounds of potassium. It is an important element, and widely distributed in nature, though always in the form of some compound, chiefly in the various feldspars.

From these, however, it has been hitherto impracticable to separate it until they have become disintegrated to clay, and until it has been taken up by plants and transformed into vegetable tissues, from the ashes of which we can dissolve potash as we have seen, and from the potash finally separate the potassium.

In the soil which we examined with a microscope, we found besides the minerals already studied a quantity of "thin, scale-like, and glistening" particles. These are flakes of *mica*, and although so different in appearance are of nearly the same chemical composition as the

feldspars. Clay, which is derived from feldspar, is a hydrous silicate of aluminum and potassium. Mica is a silicate of aluminum and potassium, with the frequent addition of small quantities of other elements.

Mica, like feldspar, is the name of a group or family of minerals consisting of several species. The most important characteristics of this group are its remarkable *cleavage*, parallel with the basal planes of the crystals, the wonderful *thinness* of the plates into which it splits—sometimes no thicker than the one 300,000th part of an inch—and above all the *elasticity* of these plates, or *lamellæ*. By these peculiarities the identification of mica is rendered easy and certain.

Of the several species of mica only two need be noticed now, *muscovite* and *biotite*, specimens 26 and 27. Determine for yourself the physical properties of these specimens.

Muscovite contains a much larger percentage of silica than biotite, a much smaller proportion of iron (sometimes none), and only about one half the amount of magnesia.

The colors of muscovite are light, being white, gray, and more rarely brown and yellow; those of biotite are dark, ranging from dark green to black.

Biotite has also the peculiar property of showing two different colors when held against the light, if the specimen is turned from side to side. This property is called *dichroism*, and biotite is a *dichroic* mineral. This is not true of muscovite. After feldspar, mica is the

most common and abundant silicate, and forms a large part of such rocks as granite, gneiss, and mica-schist. Its use instead of glass in stoves, lanterns, and other places where there is great heat, is familiar. In Siberia, Peru, and Mexico, where plates a yard in diameter are found, mica is used as a substitute for glass in windows.

Mica sometimes occurs in beautiful rhombic or hexagonal crystals.

The curious fact may be worth noting that three of the substances which we have found mingled in our handful of earth, namely, chalk, charcoal, and clay, have for hundreds of years been somewhat fantastically joined together as emblematical of the characteristics of true and loyal service, representing freedom, fervency, and zeal.

'What is freer than chalk," runs the old proverb, "the slightest touch of which leaves a trace behind? What is more fervent than charcoal? for when properly heated it causes the most obdurate metals to yield. What is more zealous than clay, which adheres to everything it touches, and which as earth brings forth every living tree and herb?"

We certainly cannot contemplate without interest these most common minerals, when we consider their inestimable value to mankind; and we cannot fail to be reminded of the danger of condemning anything as worthless from its external appearance, since we have seen chalk glorified in the calcium light, and in the form of marble preserving the beautiful conceptions

of the world's most famous artists; charcoal turning night into day when set in the electric lamp, and exciting the admiration and desire of mankind when crystallized in the diamond; and clay yielding as if by magic now the bright and untarnished metal, aluminum, and again the brilliant and exquisitely colored crystals of the sapphire and the ruby; nor should we forget those less dazzling but more important properties of these substances by virtue of which they have contributed more than any others to the education and refinement of mankind: the chalk and clay giving to our schools and colleges their crayons and slates; and the charcoal furnishing the basis of the ink which renders writing and printing possible, and in the form of graphite giving us the lead-pencil, which is even more powerful than the pen. These products are of far more value than metals or gems, for "the price of wisdom is above rubies."

QUESTIONS ON CHAPTER 14.

1. Describe potassium.
2. Uses of potash?
3. What is mica?
4. Describe biotite.
5. Describe muscovite.

CHAPTER XV.

A LUMP OF SALT.

In our study of water we learned that it has the power of dissolving a great many substances, and we noted particularly the fact that rivers in their progress to the sea carry the elements of salt, enough in the aggregate to make the whole ocean briny. Salt is so universally distributed that we feel almost certain that a portion of it is contained in our garden soil, although we have not as yet detected it in the specimens of earth which we have examined. Let us make a special search for it. It has long been known that many substances when strongly heated in a nearly colorless flame impart to the flame peculiar colors by which the substance heated may be recognized.

In order to test this the wick of your alcohol lamp should be clean, and the flame should give little light, being yellow only at the tip.

You can get a good alcohol flame by taking a clean wick, doubling it, and leaving the doubled part outside for the flame. Fig. 25.

If the flame becomes colored by particles of minerals falling on the wick, take the wick out and double it at a fresh place.

Take a small grain of salt between the tips of a pair

of platinum-pointed forceps, or on a small loop made in the end of a piece of platinum wire, and pass it up and down the edge of the blue part of the alcohol flame, moving it from below at the side of the wick up toward the tip. Notice the bright yellow color it imparts to the flame. This peculiar yellow flame is characteristic of one of the two elements contained in common salt, namely *sodium*, and is proof of the presence of that element in some form in the flame. Now, salt is one of the most common compounds of sodium, and if our specimen of soil causes this yellow color to appear in the flame we may fairly infer that it contains more or less salt.

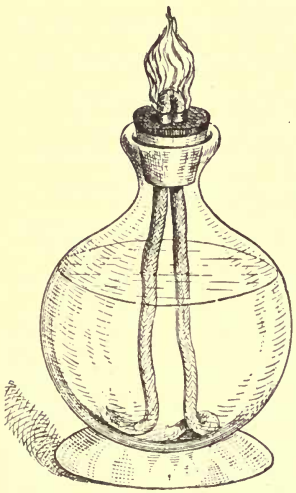

Fig 25.

Try it. Clean your forceps or wire and hold a little of the earth in the flame just as you held the salt. Do you see the same yellow color?

We shall not spend much time in studying salt as such, for our main purpose just now is to get at the elements of which it is composed. We will notice, however, its familiar and peculiar taste, its whiteness, its fusibility, and the great ease with which it dissolves in water. It crystallizes in transparent cubes, as may be

seen by slowly evaporating a gently heated solution. If the brine is heated nearly to the boiling point, and rapidly evaporated by the exposure of a large surface to the air, the crystals are very small, as in fine table-salt. The cubical form even of these little crystals may be seen under the microscope, although the crystals of table-salt commonly have their edges rounded like pebbles, partly from rubbing against one another, and partly from being partially dissolved by moisture absorbed from the air. You should therefore prepare fresh crystals for examination.

The quantity of salt used for food, either directly as a condiment, or indirectly in the preservation of meats and fish, and for other purposes is enormous. It has been estimated at fifty pounds a year for each person in the United States, and twenty-two pounds for each person in Great Britain. The coarser qualities of salt are mostly made from sea-water or other brines by a natural process of evaporation. The finer grades result from heating the brines artificially.

Salt is also found occurring in a solid form, both massive and crystallized, and is then known as rock salt. It is white, gray, or more rarely, owing to the presence of impurities, red, violet, and blue. In some places it occurs in mountainous masses. A hill of rock-salt in Spain is 500 feet high. The island of Ormuz in the Persian Gulf is formed of rock-salt. The Indus river in the upper part of its course forces its way through hills of rock-salt, which rise in cliffs 100 feet above the water. Salt is also found in beds deep under ground,

and is obtained by mining. No less than twenty-three of the United States are engaged in the production of salt. New York and Michigan are the most productive.

As an essential part of the food of live-stock salt is a necessity upon every farm. Instances are not uncommon where beasts of burden have died from lack of it.

Milch cows need a daily allowance in order to preserve the sweetness of their milk, and the quality of the fleece of sheep depends to a great extent upon their having a sufficient supply of salt.

In freezing-mixtures salt is of great use, large quantities being used in the manufacture of ice-cream.

As salt is the source from which soda is derived, it is really the basis in the manufacture of soaps and glass. Other chemical products are the chlorine used in the bleacheries and hydrochloric or muriatic acid.

The processes for the production of the metal sodium from salt are too complicated to be detailed here, but they consist in first decomposing the salt by means of heating it with sulphuric acid, which results in the separation of hydrochloric acid, and *sodium sulphate.* Afterward the sodium sulphate, or "salt-cake," as it is called, is heated with powdered coal and reduced to *sodium sulphide:* and the sodium sulphide is converted into *sodium* carbonate by means of chalk or limestone with which it is heated in a furnace of special construction. The next operation consists in freeing the carbonate of sodium from its impurities, which is accomplished by dissolving it out in water, and then evaporating the water, and calcining the residue.

The carbonate thus obtained is popularly known as "soda-ash," and is manufactured by the hundred thousand tons every year for use in glass-making, soap-making, bleaching, etc. From this "soda-ash" the metal *sodium* may be obtained by decomposing it with an electric current, or by reducing it with carbon.

Sodium is a very light metal, its specific gravity being 0.97, or very nearly that of ice. It floats upon water, and owing to its affinity for oxygen rapidly decomposes the water, disengaging the hydrogen. If the water be hot, or be thickened with starch, the globule of metal becomes so hot as to ignite the escaping hydrogen.

It will be worth your while to procure a small piece of sodium, not only that you may become familiar with so important an element in its purity, but also that you may observe its action upon water. Sodium is a silver-white metal, soft at ordinary temperatures, and melting readily. It is so widely distributed in nature that there is not a speck of dust entirely free from its presence.

QUESTIONS ON CHAPTER 15.

1. How can we detect the presence of a minute portion of salt in the soil?
2. How should the wick of an alcohol lamp be arranged so as to yield a nearly colorless flame?
3. What color does salt or sodium impart to flame?
4. Describe the physical properties of common salt.
5. What is the form of its crystals?
6. How and where is it obtained?
7. Its uses?

8. Give an outline of the process for procuring sodium from salt.

9. Uses of carbonate of soda?

10. Describe the metal sodium.

11. Its effect upon water.

12. Its distribution.

CHAPTER XVI.

MURIATIC ACID.

In the preceding chapter we incidentally learned that the first step in the process of separating the metal sodium from common salt is the addition of sulphuric acid to the salt; and we further learned that this acid causes the formation of hydrochloric acid and sodium sulphate. We then neglected the hydrochloric acid and followed the further treatment of the sodium sulphate, from which we finally obtained the metal *sodium*.

Let us now consider the hydrochloric acid. In the first place we must make enough to study. This is a very simple matter. Pour half a teaspoonful of strong sulphuric acid on a few grains of common salt in a test-tube. What takes place? What is the appearance of the gas that is given off? This gas is hydrochloric acid; the substance remaining in the tube is sodium sulphate.

Hydrochloric acid in the form of gas is very readily absorbed by or dissolved in water, just as ammonia gas is, and the hydrochloric, or "muriatic," acid of the drug stores is merely water saturated with this gas.

The popular name, "muriatic," is derived from the Latin, *muria*, meaning *brine*, and refers to its production from salt.

Make the following experiment, and you will obtain

a plentiful supply, both as a gas and in the liquid solution.

Arrange apparatus as in Fig. 26.

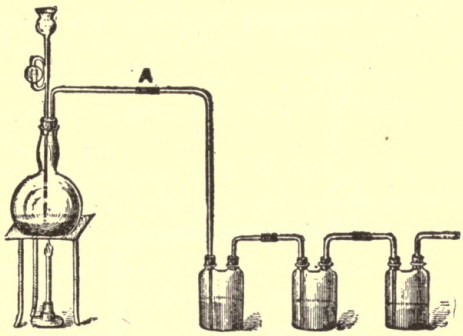

Fig. 26.

Weigh out separately 100 grains of common salt, 100 grains of concentrated sulphuric acid, and one part of water. Mix the acid and water, taking the precaution already noted of pouring the acid in very slowly while the mixture is constantly stirred. Let the mixture cool and then pour it upon the salt in the flask. Now heat the flask gently over an alcohol lamp or a gas-jet, and the gas will be regularly given off. Conduct it at first through two or three double-necked, or "Wolff's bottles," until what passes over is completely absorbed in the first Wolff's bottle. This is to get rid of the air. At first a bubbling occurs in all the bottles as the air in the apparatus is being driven out. When the air has all been expelled, the bubbling ceases and the hydrochloric gas is all absorbed in the water in the first bottle.

After the gas has passed for ten or fifteen minutes, you may disconnect the tube at A. Notice the fumes. Breathe upon them and observe that they become more dense. Apply a lighted match to the escaping gas.

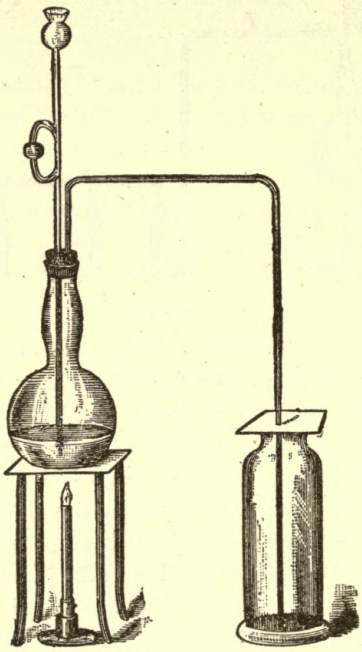

Fig. 27.

Does it burn? Collect some of the gas in a dry jar or cylinder by letting the delivery tube extend to the bottom of the jar, and covering its mouth with a piece of paper. Fig. 27.

The specific gravity of the gas being greater than that of air, the jar must, of course, have its mouth upward. Has the gas any color? Is it transparent? Insert a burning taper in the jar. Does the gas support combustion?

Now, connect the generating-flask again with the bottles containing water, and let the process continue until no more gas comes over. What substance remains in the flask?

You now have in your first Wolff's bottle ordinary hydrochloric acid in solution, and you may test it as follows: Put a little granulated or chipped zinc into a test-tube and pour some of the solution upon it.

By applying a match to the resulting gas, you will recognize it as hydrogen, which, you recollect, we have previously obtained by the action of hydrochloric acid upon zinc.

Add ten or twelve drops of the solution to a teaspoonful of water in a glass. Taste the dilute solution. How would you describe the taste?

Dip a piece of blue litmus-paper into the solution. What is the effect?

These tests are sufficient to indicate that the solution contains hydrochloric acid, HCl, which is a compound of hydrogen and a new element called *chlorine*. This element we will next endeavor to separate from the hydrochloric acid for our examination.

QUESTIONS ON CHAPTER 16.

1. How can hydrochloric acid be obtained from common salt?

2. What substance is left behind?

3. How do we obtain hydrochloric acid in a liquid form?

4. What is a "Wolff's bottle?"

5. How may we recognize hydrochloric gas?

6. Describe it.

7. What is its composition?

8. What is the derivation and meaning of the word, "muriatic?"

CHAPTER XVII.

CHLORINE.

In order to separate the chlorine, Cl, from hydrochloric acid, HCl, it is obviously necessary merely to remove the hydrogen from it. This can readily be done by adding to the acid some substance, such as oxygen, which has a stronger attraction for the hydrogen, than the chlorine has. The most convenient substance in practice is manganese dioxide, which we have already handled when we were preparing oxygen. When brought together with hydrochloric acid, the manganese dioxide parts readily with its oxygen, the oxygen unites with the hydrogen, and part of the chlorine is set free. The reaction is represented thus:

$$MnO_2 + 4HCl = MnCl_2 + 2H_2O + 2Cl.$$

This method is too expensive to be used in the manufacture of chlorine in the immense quantities required for commerce, and cheaper processes have been devised, with which, however, we are not now concerned. We will, therefore, use the manganese dioxide as follows: Put into a flask three or four ounces of black oxide of manganese. Pour upon it enough ordinary concentrated hydrochloric acid to cover it completely. Arrange the apparatus as shown in Fig. 27, on page

158, except that in order to secure a very gentle heat, the flask may be set in a pan of sand over the flame.

Heat very gently, and collect six or eight dry cylinders or bottles full of chlorine in the same manner as described in the preceding chapter for collecting hydrochloric gas.

You can see when the vessels are full by the color of the gas. The oxide of manganese used for this experiment should be in the form of small lumps free from powder. This experiment, like all others concerned with chlorine, should be carried on in a place well ventilated by a strong draft, as by an open fire-place, under a metal hood connected with a chimney, or out of doors. *Do not breathe any of the gas.*

You may now proceed to study the chlorine you have collected. You observe that it is greenish-yellow. Notwithstanding the precaution of having a good draught, you will probably perceive that it has a disagreeable odor, and that it irritates the passages of the throat and nose. The feeling produced is like that caused by a cold in the head. Inhaled in concentrated form it would cause death. It is more than twice as heavy as air.

Into one of the jars of chlorine introduce a little finely powdered antimony. The two elements at once combine with light and heat, somewhat as iron burns with oxygen.

Into a second jar introduce a few pieces of heated copper foil. They burn brilliantly, and form a copper chloride.

Into a third jar drop a piece of paper written on with ink, some flowers, and some pieces of colored calico, all moistened. Most of the colors will be destroyed.

In the fourth jar drop a *dry* piece of the same calico. It is not bleached.

This bleaching power of chlorine is of enormous importance in the cotton and paper trades. It does not, as a rule, act upon mineral colors, nor the black tints produced by carbon. Test this by placing a piece of printed paper in a jar of chlorine. The print is not affected.

The bleaching properties of chlorine depend upon its power of combining with the hydrogen of water and liberating the oxygen, which then oxidises or burns up the coloring matters.

Besides this property chlorine is also useful as a *disinfectant*, and is largely used in destroying bad odors and the poisonous germs of disease, although this effect is probably produced also by liberated oxygen, rather than by direct action of the chlorine.

Chlorine can be combined directly with hydrogen to produce hydrochloric acid. If equal volumes of these two gases, hydrogen and chlorine, be mixed, no combination occurs so long as the mixture remains in the dark and at the ordinary temperature; but if the mixed gases be exposed to sunlight, or if a flame be applied, or an electric spark passed through them, a sudden combination occurs, and the heat evolved produces a violent explosion.

Copper, mercury, tin, platinum, silver, iron, and certain other metals unite very readily with chlorine forming chlorides. The well-known substance, "chloride of lime," is formed by passing chlorine into slaked lime.

QUESTIONS ON CHAPTER 17.

1. How can hydrochloric acid be analyzed into chlorine and hydrogen?
2. What precautions must be taken in dealing with chlorine?
3. Describe the gas, its color, smell, and effects.
4. What is the effect of introducing powdered antimony into chlorine?
5. How does heated copper foil act with chlorine?
6. What effect has chlorine upon colored fabrics?
7. Explain the reason for the bleaching power of chlorine.
8. What colors are not affected by it?
9. Explain its power as a disinfectant.
10. What occurs when chlorine and hydrogen are mixed?
11. What is chloride of lime?

CHAPTER XVIII.

IRON.

We have now accounted for everything we observed in our handful of soil, but before leaving it let us look again. Take another handful and wash it carefully in a bowl of water, pouring off the muddy water and adding clear water until we obtain clean sand. Unless your soil differs from mine you will be able to find among the grains of quartz which make up most of this sand, other grains of a black color, some of which are attracted by a magnet. These grains are a kind of *iron* ore.

When clay or soil is burned, as in making brick, the red color produced is an evidence of the presence of iron. Compare the red color of bricks with the red rust seen on iron which has been strongly heated. You may also remember that in speaking of mica, I told you that besides silica, aluminum, and potassium, it usually contains a little iron. It is to a trace of iron also that we must attribute the brown color which we found tingeing some of the grains of quartz which we examined under the microscope.

Before reading further call to mind all that you already know about this element. By its lustre and hardness, and malleability, you recognize it as a metal. You

distinguish it easily from most other metals, such as gold, silver, lead, and tin, by its dark color, its hardness, and the reddish rust or oxide which so quickly gathers on its surface when exposed to the weather.

You are familiar with the properties which make iron by far the most useful metal in the world. You know that it can be had in three common forms, wrought-iron, cast-iron, and steel; that wrought or *worked* iron is the softest form and is nearly pure iron. It is used when we require strength and toughness without stiffness or elasticity; as for horse shoes, nails, wire, chains, boilers, and bridges.

Steel, you know, is the hardest form of iron. It is not so pure as wrought-iron, but contains a certain admixture of carbon, though not so much as there is in cast-iron.

It is used when to the greatest strength we would add hardness and elasticity, and is peculiarly suitable for tools, springs, rails, wire, and the finer parts of machinery. What could take the place of steel in the manufacture of needles?

Cast-iron is less pure than steel, and is brittle. It is used where rigidity of form, without elasticity or flexibility or malleability, is desired. It melts at less heat than wrought-iron or steel, and can readily be poured into moulds. It is used for stoves, furnaces, cooking-utensils, tubing, and the framework of machinery.

One of the most valuable properties of iron is that at a red heat it becomes soft like wax, in which condition two pieces strongly pressed together unite, and become

one continuous piece. This process is called *welding*.

Another property of iron, and one by which it can be instantly detected, is its sensitiveness to the attraction of a magnet, and its power of becoming magnetic during the passage of an electric current. A bar of steel becomes permanently magnetised when it is rubbed by a magnet. To this property of steel we owe the "mariner's compass," which guides our ships and steamers across the sea; while to the temporary magnetism produced in soft iron by the passage of a current of electricity, we owe, among other things, the telegraph, the telephone, and the "*dynamo*," or electric engine, which gives us the electric motor and the electric light.

Pure metallic iron is not found in nature, except in small quantities in Greenland in a rock supposed to have been thrown out of the earth by a volcano; and in meteorites. It is a white metal with a strong lustre. Iron is usually found in combination with oxygen, with oxygen and carbon, or with sulphur.

The three principal oxides of iron (specimens Nos. 10, 11, and 12) are known as limonite, hematite, and magnetite. The history of the formation of these ores through the agency of vegetable matter is most interesting, but will more properly be given when we come to the study of geology. They are easily distinguished from one another by the colors of their powders, and by the magnetism of magnetite, and they may be known from all other common rocks by their high specific gravity. Magnetite is the richest in iron, and limonite the poorest.

The carbonate of iron is called *siderite*, $FeCO_3$ Its crystallization and cleavage are essentially like calcite. Test for yourself its other physical properties.

Iron combined with sulphur forms iron sulphide, FeS_2, known by mineralogists as *iron pyrites*, or *pyrite*. (Specimen No. 29.) It is usually found in brilliant yellow cubical crystals, although there is also a twelve-sided form, and is frequently mistaken by the ignorant for gold. It has therefore been called "Fool's gold." It may easily be distinguished from gold by its brittleness and hardness, which is so great—nearly 7—that it will strike fire with steel. Hence the name, pyrite or fire-stone.

When heated in a closed tube sulphur is given off. Verify this statement by experiment, and test the dark colored residue by a magnet.

The extraction of iron from its ores consists in first converting the ore into an oxide of iron, unless the oxides themselves are used. This is accomplished by *roasting* them. If sulphides are roasted the sulphur passes off as sulphur dioxide, and iron oxide remains.

The next step is the reduction of the oxides by heating them with carbon in the form of charcoal or coke. For this purpose the ore and carbon are mixed and heated in a blast-furnace, constructed of fire-brick and masonry. Alternate layers of ore and fuel are introduced. Limestone or quicklime is also added as *a flux*. The object of this is to form a fusible substance with the earthy part of the ore. As the heat increases the ore and flux melt together, but the iron being the

heavier liquid sinks to the bottom, whence it is drawn off into proper receptacles, after which the melted flux with its impurities is allowed to escape. This refuse is called *slag*.

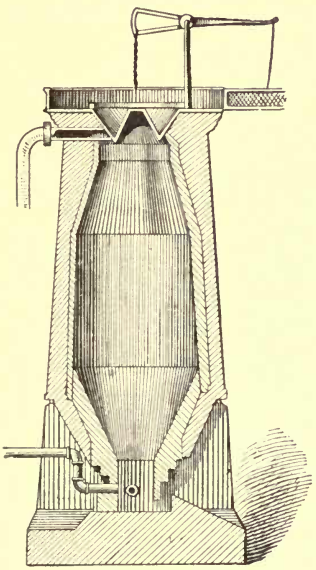

Fig. 27½.

The iron drawn off from the furnace is called *pig-iron*. It is impure, containing phosphorus, sulphur, silicon, and from 2 to 6 per cent. of carbon. It is used for *cast-iron*. Cast-iron is converted into wrought-iron in one of two ways: (1) By melting it and blowing air into the molten mass. The carbon, phosphorus, and silicon are thus oxidized and removed. This process is known as *puddling*.

(2) By mixing cast-iron with some of the purer ores, and heating to a high temperature, when the carbon, phosphorus, and silicon are oxidized by the oxygen of the ores. This process is called *cementation*.

Steel contains from 1 to 2 per cent. of carbon. There are two ways of making steel:

(1) Wrought-iron is heated with charcoal, or with other iron containing carbon.

(2) Cast-iron is melted in a large vessel, and is then partly oxidized by currents of air forced into the mass. Cast-iron is now added, and steel containing any desired proportion of carbon is thus made. This is known as the Bessemer process.

QUESTIONS ON CHAPTER 18.

1. Describe cast-iron.
2. Describe steel.
3. Describe wrought-iron.
4. State the uses of each.
5. What is the color of pure iron?
6. What are the three principal oxides of iron?
7. Describe siderite.
8. What is iron pyrites?
9. How may it be distinguished from gold?
10. How is iron extracted from its ores?
11. How is wrought-iron made?
12. How is steel made?

CHAPTER XIX.

BY WAY OF REVIEW.

We have now reached a point in our study of the world of matter where it will be profitable for us to rest for a time, and take a bird's-eye view of the path along which we have been advancing.

We began with a piece of ice, because that was as handy a specimen as we could think of, and one with which we were already familiar.

Having observed its physical properties, such as hardness, transparency, and weight, we were especially interested in the change which was caused in its structure by the heat of the hand, *i. e.*, its melting.

This led us to various experiments by which we studied the phenomena connected both with the melting of ice and the freezing again of water, including the peculiarities of its crystallization.

Having melted our ice we proceeded to examine the water produced thereby, and discovered first its property of dissolving other substances; incidentally learning something of capillary attraction, that combination of ordinary forces which results in lifting a liquid through minute tubes or pores. We next tried the effect of a greater degree of heat upon water, and found that at 212 degrees Fahrenheit it begins to be converted into

an invisible vapor called steam, a process which is accompanied by a violent agitation and bubbling, known as boiling, or ebullition; and we learned that on the withdrawal of heat this vapor is again condensed in the form of water. Observing that ice, water, and steam are not different substances, although apparently so unlike, we were curious to ascertain what it is that gives a substance its identity; or in other words, to learn what constitutes a simple element. We found that those substances are considered elementary which have never been resolved into simpler substances, and that for a long time it was supposed that the elements were four, namely, water, air, fire, and earth. This led us to examine water more carefully, and, on testing it with an electric current, we found that it separated into the two gases, oxygen and hydrogen. As these could not be further analyzed we concluded that they are elements.

In Chapter V, we studied the oxygen and hydrogen we had obtained from the water, and learned many important facts regarding chemical combinations, and particularly that sort of combination accompanied by heat and light which is called combustion.

In this connection, having observed that there is a strong resemblance between our ordinary fires and the flames which are produced by burning substances in oxygen, we examined a portion of the air which feeds our common fires, and found, as we had suspected, that it is largely made up of oxygen. Our curiosity to know why the oxygen of the air does not burn up everything

in the world brought us to the discovery that nitrogen, a powerful extinguisher of fire, is mingled in the air with the oxygen, and we afterward found this apparently inert nitrogen combined with other elements and forming highly explosive compounds like gunpowder and nitro-glycerine; we also found it constituting a large proportion of the food of animals.

Remembering the old notion that earth is an elementary substance, and knowing that that could not be true, since we had learned that at least three different elements, namely, oxygen, hydrogen, and nitrogen, enter into its composition, we determined to examine with some care a handful of earth from the garden. In this we found—besides the hydrogen and oxygen of water—quartz, limestone, charcoal, and clay, including its aluminum, and potassium; and about each of these, in passing, we gleaned such information as was at hand. Then, by the aid of its yellow flame, we discovered a little salt in the soil, and traced it to its elements, sodium and chlorine, and finally we found a trace of iron.

Our course has not been a strictly scientific one, perhaps, inasmuch as science busies itself largely with arranging the subjects of its consideration in groups or classes according to certain resemblances which exist between them—as, for example, when the zoologist groups all animals that have back-bones, under the name *vertebrate*—while we have taken our specimens in the order suggested by convenience and natural association. A science, too, walks within well-defined

bounds, closing its eyes for the time to everything outside its own narrow lines, in order that it may be more certain of seeing clearly and comprehensively everything within those lines; while we have not hesitated to leap over the fence every now and then, whenever anything of interest appeared in the fields that bordered our pathway. We have been taking a ramble along the brooks, and through the fields, rather than working methodically in a laboratory or museum.

It will be no hindrance, however, to our progress in science that we have thus pleasantly gathered and studied, even though apparently at random, a goodly collection of mineral specimens, before undertaking to make any systematic classification.

In fact, there seems to be something absurd, or at least unnatural, in that method of studying natural science which begins by committing to memory the classified results of the labors of great scientists, instead of following their example, by first collecting our material and afterward learning how to arrange it.

One of the chief disadvantages of the ordinary method of acquiring a knowledge of nature by the routine of definition learning and system memorizing is that in that way we get our information in narrow and artificial fragments, and lose sight of the intimate natural relations that exist between all the sciences, and fail to comprehend that all science or knowledge is essentially one. Thus it often happens that we hear a young student, who has read a certain chemical textbook, remark that he has "finished chemistry, and is

about to take up mineralogy;" and I have even known persons who thought they had mastered geology when they had learned the names and characteristics of the several geologic strata, although they were quite ignorant of the chemical composition of the most common minerals, and knew nothing at all of plant or animal life and structure.

Our work has not been in vain if we have learned the central truth that all the various sciences are not only interdependent, but are all merely parts of one harmonious whole, so that no one can be said to have "finished" any one science until he is master of all science. Those students who are seeking for new facts with a view to increasing the world's knowledge do well to confine their investigations closely within special and narrow lines; and on this principle a "specialist" merits praise for devoting his whole life to the study, it may be, of one family of spiders, even though he may never allow himself to notice a butterfly, a bird, a crystal, a tree, a mountain, or a cloud; but the better education for the great majority of us is one that by the faithful study in all their relations of a few typical specimens leads us to a deeper and broader understanding of the whole realm of nature, including man. Only we must be sure that so far as we do go in any direction our knowledge is accurate and thorough, and based upon the evidence of our own senses, and the conclusions of our own reason. If this book were confined to the study of the physical properties of minerals merely, as their form, weight, and color, and to a con-

sideration of those changes only which are produced in them by physical forces, such as motion and heat, it might have been called a "natural philosophy" or "physics;" if it treated only of the composition and decomposition of minerals, and of those more intimate changes which are effected when different elements are brought into chemical contact, it would have been a kind of chemistry; while if it had considered mineral specimens merely as such, describing their appearance, structure, and uses, teaching how to recognize them at sight, and telling where they may be found, it would have been essentially a little work on mineralogy. As it includes, however, as much of all these varieties of knowledge as seems necessary to a clear and comprehensive understanding of a few of the more common minerals, and to show how the others should be studied, it has seemed best to give it the more modest title which it bears: A Guide to the Study of Minerals.

It is our hope, therefore, not to give our readers the unfounded notion that they have "finished mineralogy," but to show them how they may now profitably begin that important study; and if those who follow us through this book shall find themselves better able to read with pleasure and understanding the standard works on physics, chemistry, and mineralogy, and to perceive how these are all related to one another, we shall feel, as the old phrase runs, "that our labor has not been in vain!"

The proper order for your future study seems to me to depend upon the principle that the simpler or more

elementary should come before the more complex; the lower should precede the higher. In any doubtful case you may determine which of two branches of science is the lower and which the higher by applying a law first clearly stated by President Mark Hopkins, of Williams College. He calls it the "law of the conditioning and the conditioned." It is very simple, and is this:

Whatever is a necessary condition of the existence of anything else, is as a rule lower than that of which it is the condition.

For example, in the three kingdoms of nature, animal, vegetable, and mineral, the mineral kingdom is a necessary condition for the existence of the vegetable kingdom, and the vegetable kingdom is a necessary condition for the existence of the animal kingdom. Therefore, the mineral kingdom is the lowest, and the animal kingdom is the highest.

Using this principle, now, to help us arrange a right course of scientific study, we shall find that the student having mastered the rudiments of language, arithmetic, and geography, should first be made familiar—as we have been by our simple observations and experiments —with the more common elements of the mineral world; at the same time learning the laws which govern their physical and chemical changes. Having studied then, physics, chemistry, and mineralogy, he will be led to observe how from the mineral or inorganic world, life first springs in the lower and higher forms of vegetable growth, and he will naturally make a study

of plants with the aid of some standard botany. The consideration of animal life will follow, including zoology, anatomy, and physiology.

A good review of the last four branches will be obtained by reading a good biology, and making the observations and experiments therein described.

Geology will follow, and while pursuing it, the student will find to his delight that all his previous work has been directly preparatory to a clear understanding of the structure of the earth.

Astronomy may well come next, and is but the extension to other worlds of the same methods of study that have given a knowledge of the earth. By this time the student will have laid a firm foundation on which he can stand when he approaches the study of the human intellect, and the supreme questions and problems involved in moral philosophy and natural theology.

QUESTIONS ON CHAPTER 19.

1. Give an outline or synopsis of the preceding chapters.

2. Explain the connection between the several natural sciences.

3. How many different sciences contribute to a complete understanding of minerals?

4. What is Hopkins' law of the conditioning and conditioned?

5. Give an illustration of its working.

6. Give the natural order of studying the sciences in accordance with this law.

CHAPTER XX.

WHAT IS A METAL?

We have now become somewhat familiar with eleven of the most common and abundant elements, namely, oxygen, hydrogen, nitrogen, chlorine, carbon, silicon, sodium, potassium, calcium, aluminum, and iron; and we have incidentally become acquainted with quite a number of their more important compounds.

It is now time to compare these elements with one another for the purpose of arranging them if possible in groups or classes.

To do this properly we must group together elements which have certain properties in common, and the more essential and important the properties are which we select as a basis for our classification, the better our work will be. Every mind, whether educated or not, is more or less constantly busy in classifying the objects with which it becomes acquainted. Children and illiterate persons commonly seize upon the most obvious peculiarities or characteristics of the objects they perceive as a means of grouping them, color and size being perhaps most frequently selected. Thus many birds and fishes are popularly named from their colors, as black-birds, and black-fish, blue-birds, and blue-fish, yellow-birds, white-fish, and the like. In the case

of minerals, this method would evidently be unwise, for as we have seen the same mineral often appears of many different colors. We might better use the "streak" as a basis of classification, but even this would fail us in the case of liquids and gases.

It might seem at first thought that an excellent plan would be to group together in one class the gases, in another the liquids, and in another the solids; but when we remember that very many substances readily pass from one of these states into another, we see that this would be unsatisfactory, unless we also indicated the *temperature.*

Probably the best and simplest general division of minerals is that which separates them into metals and non-metals.

While it might puzzle you to give a true definition of a metal, yet of the eleven elements we have studied you would have no difficulty in recognizing as metallic iron, sodium, potassium, calcium, and aluminum, and the other six as non-metallic. You would probably rest this decision chiefly upon the peculiar lustre of the metals, and upon their fine texture or *grain.*

In this case you would be right, and these characteristics are generally sufficient to enable us to recognize pure metals at sight, but in the case of compounds of the metals, such as ores, we need a more intimate knowledge of their essential properties.

The chemist's conception of a metal is determined rather by the nature of the compounds it helps to form than by its lustre and grain.

You have already met with a number of compounds called *acids*, notably hydrochloric, nitric, and sulphuric acids, and you have noted the sharp taste, and the powerful corrosive action which they possess; likewise their effect in changing the color of blue litmus to red.

You have also met with another and different kind of compounds, called *alkalies*, potash for example, whose, action is precisely opposite to that of acids, so that acids and alkalies are said to *neutralize* each other.

Now these alkalies are typical examples of a class of substances known as bases, whose properties are directly opposite to the properties of acids.

Among the more common bases are caustic soda. NaOH; caustic potash, KOH, and lime, CaO_2H_2.

You will notice that all of these—and the same is true of all the other bases—are formed by the union of a metal with oxygen and hydrogen. You will also observe that all acids contain hydrogen, but no metal.

Let us now try the effect of combining an acid with a base.

Dissolve ten grains of caustic soda in a glass of water. Add hydrochloric acid slowly, testing the solution from time to time by dipping into it a piece of blue litmus-paper. As long as the solution causes no change in the color of the paper, you may know that it remains alkaline; the instant the solution changes the blue of the paper to red, you know that it has passed the point of neutralization. When this point is reached, pour the water into a saucer and evaporate to complete dryness. Taste the substance remaining. It is com-

mon salt or sodium chloride. Is it an alkali? Is it an acid? Is it neutral?

Its formation may be represented by the following equation:

$HCl+NaOH=NaCl+H_2O$.

Notice that the hydrogen of the acid has changed places with the metal of the base, and that the result is water and salt. We may now sum up the results of our study thus:

An *acid* is a substance containing hydrogen, which it easily exchanges for a metal.

A *base* is a substance containing a metal combined with oxygen and hydrogen, and it easily exchanges its metal for hydrogen when treated with an acid.

The products of the action of an acid on a base are first water, and then a neutral substance called a *salt.* Sodium chloride, potassium chloride, sodium nitrate, potassium nitrate, and sodium sulphate are salts.

Bases being composed of metals with hydrogen and oxygen are commonly called *hydroxides* of their metals. Thus caustic soda is *sodium hydroxide.*

Every metal can form a salt with every acid.*

It was once thought that oxygen was a necessary part of every acid, and hence it obtained its name, which means *acid-former.* And although we have learned that there are exceptions to this—as in the case of hydrochloric acid, HCl—it is still believed that acid properties are generally due to oxygen.

*While this statement is theoretically true, it must not be understood that every metal has been actually combined with every acid to form a salt. Many theoretical salts are still practically unknown to the chemist.

It will now be clearer to you why it is reasonable to divide the elements into two classes, metals and non-metals; and you will also, at least partly, understand the division of compounds into acids, bases, and salts.

QUESTIONS ON CHAPTER 20.

1. Into what two classes may the elements be divided?
2. What is an acid?
3. What is a base?
4. What is a salt?
5. What effect has an acid on a base?
6. What are hydroxides?
7. Derivation of the name oxygen?

CHAPTER XXI.

FAMILIES OF ACID-FORMING ELEMENTS.

The eleven elements which we have thus far considered constitute at least ninety-nine per cent. of the whole substance of the earth. Most of the other elements are of comparatively rare occurrence.

It will be well, however, briefly to glance at the more important of them. In the preceding chapter we learned to divide the elements into metals and non-metals; recognizing the metals partly by their peculiar lustre and grain, and partly by the fact that they are *base-forming* elements in distinction from *acid-forming* elements. We might also have observed that as a rule they are insoluble, or soluble with difficulty, that they are opaque, that many of them can be beaten into thin leaves, or drawn into slender wire; and that most of them make a ringing sound when dropped on any hard substance.

Proceeding now to a further subdivision of the elements, we notice that they naturally fall into families, according to their chemical properties, the members of each family showing striking resemblances.

Taking first the acid-forming elements, by far the smaller of the two great divisions, as most of the elements are metallic, we have the following families:

Chlorine Family.	*Sulphur Family.*	*Nitrogen Family.*	*Carbon Family.*
Chlorine	Sulphur	Nitrogen	Carbon
Bromine	Selenium	Phosphorus	Silicon
Iodine	Tellurium	Arsenic	
Fluorine	Oxygen	Antimony	

It will not be necessary to go into details in dealing with these families. You have already studied one or more typical members of each of them except the sulphur family, and the other members may be treated briefly.

THE CHLORINE FAMILY.

The two members of this family which show the most marked resemblance to chlorine are bromine and iodine. Fluorine is not known in the free state. Its compounds, however, resemble the compounds of chlorine, and hence the element is generally included in this family.

Bromine, at ordinary temperatures, is a heavy dark red liquid, which is easily converted into vapor. It has an extremely disagreeable odor, as is implied in its name, which means a stench. Its properties are, in general, like those of chlorine, with which it is found associated in nature in salt-beds, and usually in the form of sodium bromide, or potassium bromide. It acts violently upon animal and vegetable substances, attacking particularly the skin and the membranes lining the passages of the throat and lungs. It must therefore be handled with great care. Its many compounds with other elements are called *bromides*.

Iodine occurs in nature in composition with sodium,

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in company with chlorine and bromine, but in smaller quantities. It also occurs plentifully in sea-plants.

In some parts of the ocean sea-weed is cultivated for its bromine. At ordinary temperatures iodine is a grayish black, crystallized solid. It melts easily and boils, forming a violet vapor.

To see this vapor, mix about a grain of potassium iodide with twice its weight of manganese dioxide. Add a little sulphuric acid in a test-tube, and heat gently. In the upper part of the tube some crystals of iodine will be deposited.

When a solution containing *free* iodine is treated with a little starch-paste the solution turns blue. Bromine and chlorine do not form blue compounds, and this is one means of distinguishing iodine from them.

Test this by grinding a few grains of starch to a paste with cold water in a saucer, and adding a cup of boiling water. After cooling add a little of this paste to a dilute water-solution of iodine.

Fluorine occurs in nature in large quantity, and widely distributed, but never alone. It is found chiefly in combination with calcium, in the form of *fluor-spar*, or *calcium fluoride* (specimen No. 4), and in combination with sodium and aluminum, as *cryolite* (specimen No. 8), a mineral which occurs abundantly in Greenland.

A very powerful acid gas is made by treating fluor-spar with sulphuric acid. It is known as *hydrofluoric* acid, greatly irritates the vocal organs, being therefore dangerous to inhale, and has the property of dissolving

glass. For this reason it must be kept in vessels of rubber, lead, or platinum.

It is used for etching on glass, particularly for making scales on thermometers and other glass instruments. If you would observe this action of hydrofluoric acid on glass, put five or six grains of powdered fluor-spar into a leaden vessel and pour over it enough concentrated sulphuric acid to make a thick paste. Cover the surface of a piece of glass with a thin layer of wax or paraffine, and scratch some letters through this down to the glass. Put the glass, waxed-side down, over the vessel containing the fluor-spar, and let it stand half a day. On removing the glass and scraping off the wax the letters will be found etched on the glass.

SULPHUR FAMILY.

Sulphur is familiar to every one from its use in the manufacture of matches, and from its increasing employment as a disinfectant. Sulphur candles are now burned in most well-regulated households, after every case of measles, scarlet fever, or other contagious or infectious disease. Sulphur occurs in nature both free and in combination with many metals. It is found free in certain volcanic countries, especially Sicily and Iceland, in transparent, yellow, rhombic, octahedral crystals, Fig. 28.

The compounds of sulphur with metals are called *sulphides*, and they constitute the ores from which the metals are usually obtained. Thus lead sulphide, or *galena* (specimen No. 30), zinc sulphide or *blende*, and

copper sulphide are the substances from which lead, zinc, and copper are generally procured. Sulphur also occurs naturally combined with metals and oxygen, and such compounds are called *sulphates*. Of these calcium sulphate or gypsum (specimen No. 2), barium sulphate or *heavy spar*, and sodium sulphate or *Glauber's* salt occur in the largest quantity. Pure sulphur, like pure water, is obtained by distillation. If the vapor of sulphur is quickly cooled below its melting point it solidifies, or "freezes," in the form of a fine crystalline powder, or "snow," called "flowers of sulphur;" in this particular also bearing a marked resemblance to water.

Fig. 28.

When sulphur is gently heated it melts, and may be cast into sticks, when it is known as *brimstone* or *roll sulphur*. Sulphur burns with a bluish flame, combining with oxygen to form sulphur dioxide (often called sulphurous acid), which is given off as a gas, having that peculiar suffocating odor which is perceived when a common match is lighted.

Sulphur is one of the most interesting and important elements, and can be obtained at trifling cost. It will well repay you to make a careful study of it. Allow melted sulphur to cool slowly, and observe the long, transparent, needle-like, prismatic crystals which are formed, so different from the natural crystals represented in Fig. 28.

Expose these transparent crystals to the air for a few days, and observe the curious changes that take place in them.

Pour melted sulphur heated to 230 degrees into cold water, and observe the soft, india-rubber-like mass that is produced. Remove this from the water and expose it for a few hours to the air. What change do you notice?

Sulphur combines directly with chlorine, carbon, and most other elements, and many metals burn in sulphur vapor as in oxygen, uniting with it to form sulphides.

Sulphuric acid, or hydrogen sulphate, H_2SO_4, is the most important and useful acid, as by its means nearly all other acids are prepared, and also because it is used in the arts and manufactures for a great variety of purposes.

It has been said that the commercial prosperity of a country may be judged with great accuracy by knowing the amount of sulphuric acid which it consumes. Among its many uses are the making of "soda," or sodium carbonate, which is necessary for the manufacture of soaps and glass; the production of phosphorus and of artificial fertilizers; and the refining of petroleum. More than a million tons of sulphuric acid are manufactured each year in the United States and Great Britain. It is the most important manufactured chemical substance.

Selenium and *tellurium* are so rare elements, and so closely resemble sulphur in their chemical characteristics, that we pass them by. There is no great advan-

tage in describing to you substances that you cannot bring under your own observation.

Oxygen we have already studied together as fully as is necessary for our present purpose. It is the most acid element, and may quite properly be classed by itself instead of being included in any "family."

THE NITROGEN FAMILY.

Having found nitrogen in the air, and studied its properties, we will glance at its allied elements, phosphorus, arsenic, and antimony. *Phosphorus* is not found naturally in its free state, but combined with oxygen and calcium—in the form of phosphates—in the seeds of plants, the bodies, and especially the bones, of animals, and in the minerals, apatite (specimen No. 5) and phosphorite. When bones are burned a white solid mass remains, which is called calcium phosphate, or phosphate of lime. Animals obtain their phosphates from plants, which in their turn draw their supply from the soil. Finally soils derive their phosphates from the gradual grinding up, or *disintegration*, of the oldest granite rocks which contain this element dispersed in small quantities.

Phosphorus is separated from bone-ash by mixing the ash with sulphuric acid, adding charcoal to the mixture, and heating it.

Phosphorus distils over, and is condensed and cast in sticks under water. It has to be kept under cold water until used, for in the air it rapidly combines with oxygen, and takes fire with very slight friction. It is a slightly

yellow, semi-transparent solid, resembling wax. In the air it gives off white fumes, and emits a pale *phosphorescent* light, visible in the dark. From this property it derives its name, *phosphorus* signifying in Greek a bringer of light. Great care must be used in handling this substance, and it should always be cut under water. Phosphorus not only combines thus readily with oxygen, but also with other elements, such as chlorine, bromine, and iodine. Bring together in a porcelain crucible, or a saucer, a little phosphorus and iodine. Direct combustion takes place, accompanied by light and heat.

If yellow phosphorus be left in the light, or heated to about 240 degrees for some hours without access of air, it undergoes a very remarkable change, and is converted into a dark and opaque substance, which differs as much from ordinary phosphorus as graphite differs from the diamond. Ordinary phosphorus, as we have seen, combines actively with oxygen; red phosphorus is inactive; ordinary phosphorus is very poisonous, its vapor even, when inhaled, producing a disease of the bones; red phosphorus is not poisonous; ordinary phosphorus is soluble in carbon disulphide; red phosphorus is not.

Ordinary friction-matches are tipped with a mixture of phosphorus, glue, and potassium chlorate, usually assisted by a little sulphur.

Arsenic occurs naturally combined with iron, copper, nickel, and other metals, and also as an oxide. Pure arsenic has a metallic lustre. When heated sufficiently

in the air it burns with a bluish flame, yielding poisonous fumes that have the odor of garlic or onions. Arsenic in its pure state is not poisonous, but becomes so when combined with oxygen.

When arsenic oxide is added to a mixture from which hydrogen is being evolved a compound of arsenic and hydrogen is formed. It is known as *arsine*, and is represented by the formula, AsH_3. This arsine, which is a very poisonous, colorless gas, is important as a means of detecting the presence of arsenic. When heated, it separates readily into arsenic and hydrogen, and if a cold object, as a piece of porcelain, is brought into the flame of burning arsine, the arsenic is deposited in the form of a dark spot. Arsine and other compounds of arsenic are so fatally poisonous that we advise you to make no experiments with them.

Antimony is a silver-white, metallic-looking substance. It occurs most frequently in combination with sulphur.

Boron is sometimes classed in the nitrogen family, to the members of which it bears many resemblances, though it has peculiarities which distinguish it from them. It occurs in nature most commonly in the form of *borax*, the salt of sodium. It can be obtained, though not easily, in the form of crystals which are nearly as hard as diamonds.

Borax is found in the salty incrustation on the shores of certain lakes in Persia and Thibet, and also in Nevada and California. It is also prepared artificially. It is soluble in water, yielding a clear solution with a sweetish taste. It is of great service in blow-pipe ex-

periments, because it unites with the various metallic oxides to form colored glasses. It is also used in the manufacture of enamels and glazes, and in the formation of "paste" or artificial gems. Its domestic uses for cleansing and for rendering "hard" water "soft" are well known.

Carbon and silicon, the two elements composing the carbon family, have been considered in previous chapters.

QUESTIONS ON CHAPTER 21.

1. Name four groups or "families" of acid-forming elements.
2. Describe the properties and uses of bromine, iodine, and fluorine.
3. What is the effect of hydrofluoric acid upon glass?
4. Describe sulphur. In what respects does it resemble water?
5. Uses of sulphuric acid? What is its chemical formula?
6. Describe the properties and uses of phosphorus. What precaution is necessary in handling it?
7. What is "red phosphorus?"
8. Describe arsenic.
9. What is arsine?
10. Describe antimony.
11. Describe boron.
12. Composition, properties, and uses of borax?

CHAPTER XXII.

FAMILIES OF METALS.

The metals or base-forming elements—that is elements whose compounds with oxygen and hydrogen neutralize acids and form salts—may be conveniently grouped as follows:

1. The potassium family, the principal members of which are potassium and sodium.
2. The calcium family, notably calcium, barium, and strontium.
3. The magnesium family: magnesium, zinc, and cadmium.
4. The silver family: silver, copper, and mercury.
5. The aluminium family, of which aluminium is the only well-known member.
6. The iron family: iron, cobalt, and nickel.
7. The manganese family, of which manganese is the only representative.
8. The chromium family, represented chiefly by chromium.
9. The bismuth family, represented only by bismuth.
10. The lead family: notably lead and tin.
11. The palladium family, consisting of three rare elements.

12. The platinum family, the principal members of which are platinum and gold.

It has been before observed there are many more metallic than non-metallic elements. It would not be wise to try to describe them all in this book; indeed, a thorough acquaintance with all the elements and their compounds can hardly be acquired in a life-time of patient study. Fortunately, however, the great bulk of the minerals of the world is composed of only a dozen different elements, and only a few different compounds of these; and the methods of study which we have followed in the case of these commonest substances are equally satisfactory in the case of all the rest.

Having shown you, therefore, how to study the properties of these most important elements, we may safely leave you to apply similar methods of observation to the others. Those that you will especially need to examine are magnesium, zinc, copper, mercury, lead, tin, silver, platinum, and gold. The art of extracting these metals from their ores is known as *metallurgy*.

You will find the following classification of metallic compounds convenient:

1. Compounds with chlorine, bromine, and iodine; or the *chlorides*, *bromides*, and *iodides*.

2. Compounds with oxygen, and oxygen and hydrogen; or the *oxides* and *hydroxides*.

3. Compounds with sulphur or the *sulphides*.

4. Compounds with nitric and nitrous acids; or the *nitrates* and *nitrites*.

5. Compounds with the acids of chlorine; or the *chlorates*, *chlorites*, etc.

6. Compounds with sulphuric and sulphurous acids; or the *sulphates and sulphites.*

7. Compounds with carbonic acid; or the *carbonates.*

8. Compounds with phosphoric acid; or the *phosphates.*

9. Compounds with silicic acid; or the *silicates.*

10. Compounds with boric acid; or the *borates.*

You will observe in the foregoing list that the names of some acids end in *ic* and the names of others, containing the same elements, in *ous.* The termination *ous* is used to indicate that *less oxygen* is present than in the corresponding acid whose name ends in *ic.* Thus sulphurous acid consists of the same elements as sulphuric acid, but contains less oxygen, as is shown by their chemical formulas, thus:

Sulphuric acid $= H^2SO_4$.

Sulphurous acid $= H^2SO_3$.

Notice also that the names of compounds of acids whose names end in *ic*, end in *ate*; while the names of compounds of acids whose names end in *ous*, end in *ite*; thus the salts derived from *nitric* acid are called *nitrates*, while the salts derived from *nitrous* acid are called *nitrites.*

EXERCISES FOR CHAPTER 22.

1. Observe a piece of *magnesium* wire. (Specimen No. 31.) Note its color and hardness. Test its malleability and flexibility. Is it elastic? Set fire to the wire and observe its splendid combustion.

2. Study the physical properties of *copper*, *zinc*, *lead*,

and *tin*, procuring if possible not only a specimen of the sheet metal, but also rods as large as a lead-pencil. Study also a piece of solder, which is a mixture of tin and lead; a piece of pewter which is also lead and tin, but containing a much larger proportion of tin; a piece of brass, which is composed of about six parts of copper and four parts of zinc; and a bit of bronze which contains ninety-five parts of copper, four of tin, and one of zinc, varying to seventy parts of copper and eight of tin.

Compare the color, specific gravity, and hardness of copper, zinc, and tin.

Which bends most easily?

Which "cries" or creaks when bent?

Write out the uses of each.

Are "tin" pans and pails made wholly of tin?

Test the fusibility of these metals.

3. Observe the properties of *mercury*.

What is its popular name?

Try to freeze it. (It becomes solid at—40 degrees F.)

How does its weight compare with that of water? Ascertain by reading or inquiry from what ore it is produced. Make a list of its uses.

4. Examine a piece of silver. Compare it with pure tin. Test it with acids and with fumes of sulphur. Its uses?

5. Study platinum and gold. Note their resemblances and differences. Which is the more easily fusible?

Test the effect of various acids upon gold and platinum. Test the effect of the fumes of sulphur upon them. Consider their malleability, and examine a piece of gold-leaf. What are the chief uses of gold and platinum?

CHAPTER XXIII.

HOW TO DETERMINE MINERALS.

When you find a new mineral one of the first questions that occurs to you is, What is it? The process of finding the name of a mineral from a study of its properties is called "determining" the mineral, and the various methods of accomplishing this determination constitute a science by themselves, a science known as *determinative mineralogy*. The simplest method of determining minerals, and a method that is sufficient for most of the common varieties, depends upon the study of their physical properties. As I have repeatedly said, there is no use at all in trying to study mineralogy unless you have the minerals in your hand, and unless you make the observations upon them yourself, and with the least possible help. For this reason the collection accompanying this book has been prepared. You have already found it useful as a means of illustrating the previous chapters. It remains to make use of it in showing you the simpler methods of determining the minerals you may hereafter obtain. For this purpose the specimens are numbered but not named. There is a key to the collection at the end of this book, but you are not to refer to it until you have done your best to determine the names of the specimens by the method

now to be described. After you have decided on these names as well as you can, you may turn to the key to satisfy yourself as to the correctness of your determinations.

We will first try only a dozen of the most easily recognized specimens. Select from the collection numbers 1, 2, 3, 5, 6, 7, 11, 29, 26, 28, 30, and 32.

Examine these specimens, noting their hardness, color, streak, etc., and write a description of each. Then by comparing your observations with the following brief descriptions, determine the name of each specimen. Of course, the descriptions here given will not follow the numbered order of the specimens.

HINTS.

The only tools needed are the streak and scratch-plate (furnished with the collection), and a pocket-knife.

In using the glass plate to test hardness or streak, do not scratch the plate all over with your specimen. A scratch or streak-line as long as this —— is sufficient.

Generally the color of the streak will be seen better if the streak-plate is put upon a piece of white paper.

Whenever you have made out a mineral, write the name on a small label and place it in the compartment with the specimen.

Sometimes the mineral to be determined forms only a part of the specimen. A portion of the rock in which the mineral occurs may be present, or some other associated mineral. In this case try to find which is the

essential mineral, and see whether you cannot determine also the others that accompany it.

Hardness.—Find out how many of the twelve specimens you can scratch with your finger-nail and put these aside; then find the softer and harder ones among these; see what minerals described below have hardness 1 and 2, and from the description try to determine which specimen is talc, which is kaolin, graphite, gypsum, etc. Label all those you have determined; the other specimens try with your knife; do not disfigure the whole specimen with scratches; a little scratch in some corner will tell the story, and the specimen will be as good as before. Put aside all those you can scratch with your knife, separating those that you can scratch easily from those that require force. The hardest of these will scratch the glass plate (the smooth side of this is to be used for receiving the scratches.) In this manner you will find the minerals that have hardness 3 to 6. The specimens that are still left will scratch the glass easily; they will have hardness 7 or 8.

Lustre.—Examine a porcelain dish, a mother-of-pearl button, a piece of white wax, white satin or silk, and a silver dollar: All these are white, but each has a different lustre; each reflects the light differently. The lustre of minerals often resembles that of glass, or wax, or silk, or pearl; these minerals are said to have a glassy (vitreous), waxy, silky, or pearly lustre.

In examining the minerals for their lustre, notice first whether the lustre is metallic or non-metallic. If the specimen does not shine like silver, or gold, or

brass, or copper, or steel, or any other metal, the lustre is non-metallic; you will judge next whether the lustre of your specimen resembles that of glass, or pearl, or resin, etc. Minerals that have no lustre, such as chalk and clay, are called dull. In some minerals the lustre looks somewhat metallic, but not clearly so; for these the term sub-metallic is used.

Transparency.—Some minerals are as clear as glass; they are transparent; some let the light through dimly, like thin China ware; they are translucent; some do not let any light through; they are opaque. Semi-transparent means half-transparent, not quite clear and not very dim; the streak-plate is a good example. In some minerals a faint light passes only through the edges; these are called sub- or semi-translucent.

STREAK.

The streak or powder of the specimen is of more importance for its determination than the color; it shows the essential color of the mineral. Most of the minerals that are white or colorless when pure, have a white streak even when they are colored. Take some clear glass and some colored glasses and grind each to a fine powder; the powder will, in all cases, be white (with probably a faint tint of color in some instances), because the amount of coloring matter used is very small. To test the streak, rub one of the edges or corners of the specimen against the ground side of the streak-plate, causing it to trace a small line on the latter, and notice the color. The minerals of the hardness of 7 and 8

will not leave any trace of their powder on the glass because they scratch it, but as they have, almost without exception, a white or grayish streak, it is not necessary to test them. Their essential color may be found, however, by breaking off a small chip and grinding it to a fine powder.

TASTE AND ODOR.

A few mineral salts can be distinguished by their taste; some minerals have a peculiar odor; in kaolin, and all rocks containing clay, the clayey or *argillaceous* odor is readily noticed after breathing upon them; in others the odor is brought out by rubbing the specimen against another stone or by breaking it with a hammer.

DESCRIPTION OF THE TWELVE MINERALS.

1. *Calcite* (*calc spar*).—Hardness, 3; streak, white; lustre, vitreous; color, generally white or light colored, but also of darker colors; transparent to opaque. The colorless, transparent variety is called Iceland spar. Marble is a crystalline variety with a sugary fracture. Limestone is a common, compact variety.

2. *Talc.*—Hardness, 1; streak, white; lustre, pearly; color, greenish to white; translucent; smooth, slippery, soapy to the touch. Commonly called soapstone, also steatite and French chalk. Foliated talc can be separated with a knife into thin scales, somewhat like mica, but mica is elastic, talc is not.

3. *Gypsum.*—Hardness, 2; streak, white; lustre,

pearly, glassy, silky or dull; color, white, gray, pink, brownish, etc. The colorless, transparent variety is called selenite; the fibrous variety, satin spar; the white fine grained variety, alabaster. Foliated gypsum resembles talc and mica, but it is not greasy to the touch like talc, and it is not split as easily as mica, and is not elastic.

4. *Mica* (*muscovite*)—Hardness, 2; streak, white; lustre, pearly, somewhat metallic; color, white to gray and brownish. Can be easily split into thin elastic leaves; used in stoves. Often wrongly called isinglass. Isinglass is a kind of glue, made from the air-bladder of the sturgeon, and used for making jelly.

5. *Fluorite* (*fluor spar*).—Hardness, 4; streak, white; lustre, vitreous; colorless or of various beautiful tints; more or less transparent. The specimen in this collection resembles calcite and quartz, but it is easily distinguished by its hardness; it will scratch calcite and will be scratched by quartz.

6. *Quartz.*—Hardness, 7; streak, white; lustre, vitreous; color, colorless (rock crystal); also white (milky quartz;) pink (rose quartz); brown or nearly black (smoky quartz); violet (amethyst); yellow (false topaz). These and some others are the glassy varieties of quartz; all these are transparent to translucent. Nearly all the light-colored pebbles found on river banks or lake shores are quartz; sand is usually composed of quartz grains.

7. *Orthoclase* (*feldspar*).—Hardness, 6; streak, white; lustre, vitreous. Compare this specimen with calcite. Although the lustre in both may be called

glassy (vitreous) there is some difference. In calcite the lustre is easily seen in all directions. In orthoclase the lustre inclines somewhat to the silky, and you have often to turn the specimen around in your fingers and look at it in this way and that before you see the lustre plainly; examine until every part has shown its lustre to you. Color, white, yellow, gray, pink, and other light tints; translucent, at least at the edges, sometimes transparent.

8. *Apatite.*—Hardness, 5; streak, white; lustre, vitreous; color, generally green, sometimes brown, yellowish or white; generally opaque, often translucent at the edges.

9. *Halite* (*rock salt*).—Hardness, 2; streak, white; lustre, vitreous; color, white, gray, red, rarely blue or violet. Taste, salty.

10. *Galenite* (*lead ore*).—Hardness, $2\frac{3}{4}$; streak, lead-gray; lustre, metallic; color, gray. Often contains a little silver.

11. *Hematite* (*iron ore*) (*variety, specular iron*).—Hardness, $5\frac{1}{2}$; streak, red; (rub on the streak-plate and blow away the shiny scales to see the streak); lustre (of this variety), metallic; color, gray. There are other varieties of hematite iron ore of different lustre and color, but the red streak is common to all.

12. *Iron pyrites* (*fool's gold*).—Hardness, 6; streak, dark greenish or gray; lustre, metallic; color, brass-yellow; sometimes contains gold. To distinguish it from gold, hammer it; gold flattens; pyrites crumbles into a dark, greenish powder.

CHAPTER XXIV.

FLAME TESTS, HEAT TESTS, AND ACID TESTS.

The properties by which you determined the names of the twelve specimens described in the preceding chapter are those which enable one to distinguish a mineral by looking at it, and testing it by means of a pocket-knife and a streak-plate. This is the simplest method of determination, and the student should try to determine by it as many specimens as possible. But in order to be certain that the determinations are correct, and also that he may learn the composition of minerals, he must learn to apply other tests with heat and with acids. These are also called *dry* and *wet* tests.

APPARATUS NEEDED.

For these tests you will need an alcohol lamp, or a Bunsen gas-burner, a forceps—preferably with platinum points—small open and closed tubes of hard glass, and a few test-tubes.

It will also be convenient, though not necessary, to have a small agate mortar and pestle; a few small porcelain dishes, such as are in dolls' "china sets," and a magnifying-glass.

HINTS.

Make all the tests yourself, getting from a friend or teacher any further suggestions you may need about the handling of the apparatus or the acids.

Small fragments chipped from each specimen will be sufficient for all the tests.

Be careful with the acids; they stain and burn.

Be neat and orderly. A noted mineralogist makes his determinations on a parlor table, and makes so little dust that a stroke of the handkerchief removes all traces of his work.

Keep your apparatus in a suitable box.

Clean and dry all apparatus after use.

In chapter XV, on salt, we have already given practical directions for testing the colors of flames, and these need not be repeated. Before studying any new minerals, however, it will be best for you to repeat the test there described for showing the yellow sodium flame, and also to make the following simple experiments with the specimens treated in the preceding chapter.

1. Break off a little chip of calcite, and giving it a few taps with a hammer, notice how it breaks into little crystals, some of them quite perfect. Take one of the larger crystals and hold it in the flame. Probably it will not show any distinct color, but will fly to pieces or "decrepitate."

The metal calcium, which is in calcite, gives the flame an orange color, but this is not always easily seen; it can be coaxed out by grinding a crystal of the calcite to powder, wetting it with a drop of water, and a drop

of hydrochloric acid, then steeping in this a few fibres of asbestus and holding them in the flame.

2. Try fluorite; it decrepitates more readily than calcite; it often shows the calcium flame more distinctly than the former. An interesting experiment with fluorite is the following: Heat an iron shovel almost to redness; take it into a dark room and throw some grains of fluorite on it; the fragments will become luminous, shining with a beautiful bluish or greenish light; this is called phosphorescence. Apatite often shows phosphorescence when heated; less frequently calcite and dolomite.

3. Try gypsum; see how it turns white under the intense heat. Wet it, heat it again, and notice the color of the flame; gypsum also contains calcium. After you have burnt the gypsum white, you have "plaster of Paris."

4. Treat a scale of mica as you did gypsum, and notice the difference.

5. Chip off a piece of galenite; notice the cleavage. Hold a very thin piece in the flame; after it is hot, hold it just outside the edge of the flame and admire the delicate blue color; does it melt? Does it smell of sulphur?

6. Try the same with iron pyrites; generally it decrepitates; if the galenite did not smell strong enough of sulphur, this will satisfy you.

7. Orthoclase after being heated should show a purple flame, because it contains potassium; if it does not, wet it with a drop of HCl, heat it again near the tip of

the flame, then hold it to the outer edge to see the color. [In many cases moistening the mineral with hydrochloric acid helps to bring out the characteristic flame. A little dish with a drop of HCl mixed with a drop of water is very handy for this purpose; moisten your splinter or your fibres of asbestus by touching the edge of the drop with it].

8. Apatite will generally give no characteristic color. Powder a little, wet with one small drop of sulphuric acid; wipe off part of the powder with a little asbestus that has first been heated in the flame (to burn away dust or sodium), then try the flame test. It will show a brownish color, turning to bluish green, which shows the presence of phosphoric acid.

WHAT CAN WE LEARN FROM THESE FLAME TESTS?

First—That certain minerals, on account of the elements which are in them, impart peculiar colors to the flame; these are due to the vapors which the elements give off when heated.

Intense yellow flame (bordering on orange) indicates that sodium is present.

Yellowish-red flame indicates calcium.

Violet or purplish flame indicates potassium.

Carmine-red flame indicates lithium.

Crimson flame indicates strontium.

Emerald-green flame indicates copper (also thallium).

Yellowish-green flame indicates barium or boracic acid.

Bluish-green flame indicates phosphorus.

Azure-blue (intense) flame indicates selenium.

Light-blue flame indicates arsenic.

Copper in combination with chlorine gives also a beautiful blue flame. Dip a piece of copper wire into HCl and hold it in the flame; notice the change from azure-blue to emerald-green as the chlorine evaporates.

The characteristic flame color of a mineral is often hidden by impurities or associated minerals. Especially sodium, if present, makes it difficult to distinguish the true color. If the flame is viewed through blue glass, the yellow sodium flame is not seen. The calcium flame is sometimes changed by the presence of barium or strontium.

Second—That some minerals melt, at least partly, even in the flame of an alcohol lamp.

Third—That some minerals, especially those with eminent (very perfect) cleavage, fly to pieces when heated.

Fourth—That some minerals emit a peculiar odor when heated, the most characteristic being:

Sulphurous odor—most minerals containing sulphur.

Garlic odor—minerals containing arsenic.

Horse-radish (decaying) odor—minerals containing selenium.

Bituminous odor—mineral hydrocarbons (coal, etc.).

HEAT TESTS WITH THE OPEN AND CLOSED TUBE.

Very small grains of the minerals are generally best for heat tests. To transfer them to the tube, a little trough made of paper or tin is the most convenient instrument.

Cut a strip of paper three or four inches long and one-eighth of an inch wide, and double it lengthwise; near the end of this trough put the fragments or powder of your mineral and insert it in the tube; this prevents the mineral from sticking to the sides of the glass. Push the paper into the tube until your mineral fragments are about half an inch from the end, then give the tube half a turn, and withdraw the paper. *See that your tubes are clean and dry inside and outside before you begin a test.*

HEAT TESTS FOR THE MINERALS DESCRIBED IN CHAPTER XXIII.

1. Put a few small pieces of gypsum into the closed tube and heat the closed end of the tube in the alcohol flame; notice drops of moisture gathering in the upper part of the tube; heat until no more moisture gathers; drive the water farther up by heating the part in which it has gathered. This will show that moisture is always driven from the hotter to the cooler part, evaporating in the former and condensing in the latter. Pour out the gypsum; it is white, and has again, as before, been changed by driving out the water.

Try the same experiment with powdered talc. The powder of limonite should be yellow or brown; notice its color after the water has been expelled. Try quartz, orthoclase, and hematite, and see whether you can drive out any water. (Minerals lying in damp places sometimes absorb some moisture).

2. Try fluorite, and notice how it decrepitates. Take

the tube, after heating the fluorite, into a dark room and notice again the phosphorescence.

3. Put a few grains of iron pyrites into the closed tube, and heat it; a yellow coating of sulphur will form on the cooler part; it is driven out of the pyrite in the form of vapor, and it condenses again above, where it is cooler; this is called forming a *sublimate*. Notice the odor given off while heated. Wet a narrow strip of blue litmus-paper, and insert it about one-fourth of an inch in the tube; heat the mineral again and notice the paper turning from blue to red; this is a sign that acid fumes are escaping. Try the same with galena, pulverized finely; it takes stronger heat to drive the sulphur off, but you will notice the odor and the acid reaction on the litmus paper. Of course, a tube in which sulphur has formed must not be taken for another experiment until all the sulphur is removed. You can drive it off by heat as you did the moisture, but it takes stronger heat and a little skillful management.

Fluorite and pyrites should also be tried in the open tube. Insert a few grains of the iron pyrites, by means of your trough, in the open tube, depositing it about half an inch from one end; hold the tube at the other end, slanting, just so that the pieces will not fall out; and hold the grains over the flame; notice again the sublimate, the odor, which will be stronger than before, and the acid reaction on the blue litmus-paper.

The difference between the test in the closed tube and the open tube is this: in the former the substance is heated without the presence of oxygen, because most

of the air is driven out of the tube by the heat; in the latter, there is a draft of air passing through the tube while the substance is heated, and the vapors can unite with the oxygen of the air. In the case of heating the iron pyrites in the closed tube, nearly all the sulphur settles on the cooler part of the tube, while in the open tube most of it unites with oxygen and leaves the tube, causing the strong odor.

WHAT CAN WE NOTICE IN THE TESTS WITH THE CLOSED AND OPEN TUBES?

The substance may decrepitate; in this case the mineral should be used in the form of powder.

It may give off an odor, generally of the same kind as that noticed in the flame test, but stronger.

Moisture may form in the tube, which indicates that the mineral contains water.

A coating, or sublimate, may form on the cool part of the tube. The following are the most important:

A yellow coating indicates *sulphur*; the odor will prove it: brown-red when hot, and reddish-yellow when cold indicate *arsenic sulphide:* brilliant black, metallic, *arsenic*, the garlic odor furnishing additional proof: black when hot, brown-red when cold, formed near the mineral after strong heating, *antimony sulphide:* dark red, *selenium* (horse-radish odor): a metallic mirror, composed of little metallic drops, *mercury* (quicksilver).

A few other coatings will be mentioned in the descriptions of the minerals which produce them.

Acid or alkaline vapors may be driven off; the first

will redden blue litmus-paper, and the second will turn red litmus blue. The papers must be moistened.

A change of color may take place, as with limonite, which changes into red hematite.

Phosphorescence is sometimes observed.

ACID TESTS.

HCl—Hydrochloric acid; H_2SO_4—Sulphuric acid.

For transferring the few drops of acid needed from the bottle to the test-tube, the dipping-tube, that is, a narrow glass tube about eight inches long, should be used; dip the tube about one inch into the acid, and close the top with your finger; lift it out and transfer it to the test-tube which you hold in your left hand, and when the lower part of the dipping-tube is near the bottom of the test-tube, remove the finger and the acid will flow out. But first put your mineral—a few grains, powdered—into the test-tube; then, if the mineral is to be tested with diluted acid, add a few drops of water, using the dipping-tube; and lastly put in the acid Never put the water into the tube after the acid, always before.

In heating the contents of the test-tube do not apply the bottom of the tube to the flame, but hold it slanting and let the upper part of the liquid be nearest the flame; else a steam-bubble, from the bottom, may throw the acid out of the tube. Heat gradually. You will find it safer to heat the acid by passing the test-tube repeatedly through the flame instead of holding it steadily in it. If the acid has a tendency to spurt, as it

will sometimes, plug the opening of the tube with a light wad of cotton, or better, asbestus. Do not turn the opening of the tube toward you. See that the outside of your test-tube is dry, else it is liable to crack. Have the lamp on a slate or tray. You can use the tube-holder with the test-tube, but experts use their fingers.

EXPERIMENTS WITH ACIDS.

After having read the foregoing directions carefully, you may begin to test your specimens.

1. Powder a fragment of calcite and transfer to the bottom of the test-tube; dip the dipping-tube an inch into water, then close the top with your finger and transfer the water into the test-tube. Do the same with HCl. Do not let the liquids flow down the side of the test-tube, but deposit them just above the powdered mineral. Notice the *effervescence* or bubbling up of gas. Put a drop of acid on the piece of calcite and see again the effervescence; if you have a piece of marble, try it on that. See whether the acid will have the same effect on any other mineral in your collection. Thus we have an easy method for determining calcite or limestone. After the calcite is all dissolved, heat the solution to concentrate it; then put a drop on the smooth side of your streak-plate, or any other piece of glass; also put a drop of pure HCl on the glass, and let both dry; examine both drops; of the HCl very little is left; of the calcium chloride, which was in the first drop, you may see the crystals with a magnifying glass. The glass plate should be slightly warmed.

2. Powder a little fluorite and put it into a test-tube; examine the tube first to see whether it is perfectly clean and clear, especially near the bottom.

Put in a few drops of H_2SO_4, not diluted with water; heat slowly and carefully; do not smell the fumes in this test; they are poisonous. When the acid is nearly evaporated, let the tube cool, then wash it out with water and let it dry; you will find that the glass has been attacked (corroded) by the fluorite.

Another test: Put a little powdered fluorite on the smooth side of your streak-plate; let one drop of sulphuric acid fall on it; put the plate away in a safe place over night, wash away the acid and powder and notice the effect on the glass. These are tests to detect the presence of fluorine in a mineral.

WHAT HELP CAN WE GET FROM THE ACID TESTS TO DETERMINE MINERALS?

Some minerals dissolve with effervescence in diluted HCl, especially the carbonates, as carbonate of calcium (limestone), carbonate of iron (siderite), and others.

Some minerals dissolve without effervescence.

Some minerals dissolve partly or sparingly.

The substance formed by the solution may be found by evaporating the liquid. It will generally crystallize; the crystals cannot always be seen by the naked eye, but the student possessing a microscope will find a source of great delight in their examination.

CHAPTER XXV.

TWELVE OTHER MINERALS.

We will now examine twelve other common minerals. (Specimens Nos. 8, 10, 16, 17, 18, 21, 22, 25, 27, 34, 35, and 36.)

Determine these if possible by an examination of their physical properties. Afterward use the flame, heat, and acid tests to identify any not recognized by their external characteristics, and to verify your decisions with regard to the others.

It is a good plan to enter in a blank book the observations made on each specimen as it is tested. Compare your descriptions with those that follow.

Abbreviations: H.—hardness; fl.—flame test; c.t.—closed tube; o.t.—open tube; comp.—composition; b.b.—before the blowpipe.

Italics indicate important characteristics which will help you distinguish a mineral from others that more or less closely resemble it.

DESCRIPTION OF THE SECOND DOZEN MINERALS.

1. *Sulphur.*—H., 2; lustre, resinous; color, yellow to gray or brown; streak, yellow; translucent to opaque; generally massive, sometimes crystallized, also stalactitic; fracture, uneven; very brittle; light.

Comp.—S.,—Fl., *burns* with a *blue* flame and sulphur-

ous odor, and melts; c. t., melts, then vaporizes, the vapor forming a coating on the cooler part of the tube. When the mineral is held in the warm hand a crackling noise is heard. It becomes electric when rubbed.

2. *Stibnite* (*gray antimony*).—H., 2.5; lustre, metallic; color, lead-gray; streak, gray; opaque; crystalline; columnar or fibrous to massive granular; cleavage, perfect to indistinct; brittle; rather heavy.

Comp.—Sb_2S_3 (antimony sulphide).—Fl., bluish, decrepitates, *melts*, sulphur odor; c. t., decrepitates, melts, acid reactions on litmus paper; o. t., fills the tube with *white fumes*, characteristic of ores of antimony and which leave a white coating; peculiar odor in which sulphur can readily be detected. As the fragments decrepitate violently, they should be finely powdered for this test. HCl, dil.; generates sulphuretted hydrogen, marked by its disagreeable odor. Stibnite is named from stibium, the scientific name for the element antimony. Stibnite is easily recognized on account of its melting so readily.

3. *Sphalerite* (*zinc blende*, *Black Jack*).—H., nearly 4; lustre, *resinous* or metallic; color, yellow to red, brown, and grayish; streak, yellowish white; translucent to opaque; crystalline, sometimes granular to nearly compact; cleavage, perfect; brittle; rather heavy.

Comp.—ZnS, (zinc sulphide).—Fl., does not melt; sulphur odor; c. t., acid reaction. HCl evolves *sulphuretted hydrogen*, especially when heated. Sphalerite comes from a Greek word meaning treacherous.

The German word blende means deceiving; it is often found together with galena, but yielding no lead, the miners thought a trick had been played on them by the goblins. It is now an important source of zinc.

4. *Cryolite.*—H., 3; lustre, vitreous; colorless to white and brownish; streak, white; transparent to translucent; usually massive; cleavage not always distinct; brittle, light.

Comp.—$Na_6Al_2F_{12}$; Fl., yellow soda flame; *melts* in the hot part of the flame. This test affords an easy distinction between this mineral and others that resemble it in color and lustre; take a small splinter for the test. C. t. decrepitates; acid reaction; etches the glass a little above the point where it is heated. Try the tests for fluorine. Cryolite is used for making soda, the metal aluminum, and a white glass. Siderite frequently accompanies cryolite. Cryolite means ice-stone; this name is given to it because it melts easily.

5. *Limonite.*—H., about 5; streak, yellow to brown; lustre sub-metallic to silky; color, yellow or brown; brittle; cleavage, uneven, splintery.

Comp.—$2\,Fe_2O_3 + 3\,H_2O$; c. t. yields water which condenses in upper part of tube. Compare the streak of limonite with the streak obtained by rubbing a rusted bar of iron on paper. It is reasonable to conclude from the close similarity of the streaks that we have in this limonite iron-rust or, what is the same thing, iron oxide —iron combined with oxygen taken from the air. The name limonite is derived from a word meaning a meadow or swamp, and is used because this "bog-iron-ore" is

formed in wet places. Test the streak of limonite after it has been heated in the closed tube: it is now red, not yellow. Compare it with the streak of hematite; it has the same appearance. In fact hematite is exactly like limonite from which the water has been expelled by heat.

Neither ore is attracted by a magnet. Limonite and hematite are the two great natural coloring agents. Almost all the yellow, brown, and red colors in rocks and soils are due to them.

6. *Chlorite.*—H., about 2; lustre, pearly to dull; streak, light green; color, dark green; clay odor noticed when the specimen is breathed upon; generally somewhat greasy to the touch.

Comp.—Chlorite is properly the name of a group of highly basic minerals of variable composition, but all essentially hydrous silicates of aluminum, magnesium and iron. The specimen in this collection is *prochlorite*, one of the most abundant species. It yields water in the closed tube.

7. *Celestite.*—H., 4; lustre, vitreous to pearly; colorless, white, generally with a bluish tint, also yellowish; streak, white; transparent to translucent; crystalline, brittle.

Comp.—$SrSO_4$, (strontium sulphate): Fl. decrepitates slightly. The characteristic flame is a deep *crimson red*; but it is difficult to get it in the alcohol flame without the use of the blow-pipe; you will succeed, however, in coaxing it out by treating it in the following manner:

Grind a small fragment to a fine powder; take some fibres of asbestos, dip them in the powder and hold them with the forceps in the edge of the blue flame; a yellowish flame generally appears at first, but after continued heating, especially by holding it for a little while in the hottest part of the flame, the color will turn to a reddish hue; now slightly moisten the tip of your asbestos with HCl (by holding it just at the edge of a drop of the acid) introduce into the flame again and notice the crimson flashes. Keep it in the flame for a little while; if the color is not distinct, repeat the test.

Celestite is derived from the latin word for sky, because the mineral is sometimes blue; preparations made from this mineral are used for the crimson lights in fireworks.

8. *Kaolinite.*—H., 1 to 2; color, white when pure, (then used to make porcelain) but generally yellowish or gray; streak, same as color; lustre, none, *i.e.* dull; friable, (crumbling between the finger). A yellowish clay.

Comp.—$Al_2Si_2O_7 + 2H_2o$. Infusible and insolubie. C.t. yields water.

9. *Hornblende.* — (*amphibole*).— H., $5\frac{1}{2}$; streak, white; color, greenish to black; lustre, silky to vitreous, reminding one of the lustre of feldspar; brittle, cleavage, prismatic, perfect.

Comp.—(Ca Mg Al Fe)Si O_3. Insoluble, infusible.

10. *Labradorite.*—H., 6; lustre, vitreous to pearly, feldspar lustre; color, generally light and dark gray mixed; streak, white; translucent; generally massive;

cleavage, perfect to interrupted; brittle. Labradorite often shows a beautiful *play of colors* if viewed at a certain angle, (especially in the dark gray portions); blue and green are the more common colors; yellow and red are rare. The color is generally seen best on wetting the specimen. Labradorite is a variety of feldspar (soda-lime feldspar); it is generally darker and a little heavier than the other varieties. It is named from Labrador, where fine specimens have been obtained.

Comp.—$(Na_2Ca)Al_2Si_3O_{10}$. Insoluble.

11. *Biotite*, (*black mica*).—H., nearly 3; lustre, pearly to sub-metallic; color, *black*, or nearly so; streak, gray; translucent to opaque; crystalline, foliated; cleavage, perfect, sectile. The thin leaves into which it splits are *elastic*.

Comp.—$(K\,FeMgAl)_2SiO_4$, (potassium-iron-magnesium-aluminum silicate): Fl., the purple potash flame is often noticed. Put a few small scales into a test tube, add about 10 drops of H_2SO_4 (undiluted), and heat carefully; the mineral will partly dissolve; pour off the acid, leaving what remains of the mineral in the tube, fill the tube with water to wash out the remaining acid and examine what is left of the scales; they have changed from black to white and now consist simply of the silica, which remains undissolved. Muscovite does not dissolve in H_2SO_4. Between Muscovite and Biotite there is another variety of mica called Phlogopite. (*i. e.*, having a fiery appearance); it is generally of a bronze lustre; it also dissolves in H_2SO_4. Biotite is named after the French mineralogist Biot.

12. *Chrysotile.*—H., 3.75; lustre, silky; generally greenish, but also light gray, yellow to brown; streak, white; translucent; crystalline, fibrous, separates into *fine, flexible fibres.*

Comp.—$Mg_2SiO_4+H_2O$, (hydrous magnesium silicate): Fl. does not melt, but becomes brittle, c. t. yields *considerable water.* Chrysotile is a fibrous variety of serpentine; it is commonly called asbestos, and, indeed, it resembles this mineral very closely, and is used for the same purposes for which asbestos is used, that is, for making fire-proof cloth, paper and ropes, coverings for boilers, etc. Very fine silky varieties of asbestos and chrysotile are called amianthus. The name, chrysotile, means golden fibre. Asbestos is a variety of amphibole, and the name means indestructible. Its composition is generally (Ca Mg Fe) Si O_3.

WHAT THE TESTS MEAN.

Compounds are, as a rule, so different from the elements of which they consist, that it is often impossible to tell, by looking at a mineral, of what it is composed. Therefore, some means are resorted to to separate or decompose the compound, so that we can detect some well-known and prominent property of the elements. This examination of a substance is called *analysis.* If we wish simply to know whether a certain element is present in the compound we make a partial analysis; if we wish to know every element that is present, we must make a qualitative analysis, and if we wish to find exactly how much (what percentage) of one or all the

elements is present, we must make a quantitative analysis.

ANALYSIS IN THE DRY WAY.

Flame tests.—Heat decomposes nearly all compounds. In the flame tests, part of the mineral becomes decomposed and vaporizes, and those elements whose burning vapor has a characteristic color or odor, as the yellow of sodium, the orange of calcium, the emerald green of copper, the odor of sulphur, arsenic, etc., can be detected.

The tests in the tubes.—Heat is used also in these tests to decompose the compounds; sometimes we find the elements separated in different parts of the tube, as we noticed in testing iron pyrites, when sulphur and iron were separated. Sometimes the acid-element is driven off, and reveals itself by the smell, and by the change of the color of litmus-paper; or, in the case of an oxide, the oxygen may be driven off. In heating limonite, the water is expelled, changing the hydrous yellow substance to anhydrous red hematite.

EXAMINATION IN THE WET WAY.

The chemical power which holds the different elements in a compound together, becomes weaker also through solution. Few minerals dissolve readily in water, but quite a number in HCl, H_2SO_4 or HNO_3 (nitric acid). When a mineral dissolves in an acid it unites with a part of the acid and forms a new compound. Zinc, for instance, dissolves in chlorhydric acid, and becomes

chloride of zinc. If there is already an acid element in the mineral, a stronger acid-element will drive it out and take its place. Calcite, $CaCO_3$ is calcium combined with the CO_3 part of carbonic acid, H_2CO_3; HCl is stronger than carbonic acid, it drives out the CO_3 and combines with the calcium, forming, instead of a carbonate, a chloride of calcium. The chemist puts it down in this way:

$$CaCO_3+2HCl=CaCl_2+CO_2+H_2O.$$

What becomes of the CO_3? CO_2 escapes in bubbles, causing the effervescence. The other O unites with the 2H left from the 2HCl and forms H_2O (water) which remains in the test tube.

Calcium chloride is soluble in water, like common salt. If we wish to get it in its solid form we must evaporate the solution.

In applying the acid tests the following facts are worth remembering:

An acid may form, with the mineral, a soluble compound. Then the solution is clear; there is no sediment left after all is dissolved, as we saw in dissolving calcite in HCl.

An acid may form an insoluble compound; a precipitate will then be formed, as in treating calcite or calcium chloride with H_2SO_4.

An acid may dissolve a part of a mineral, as in the case of biotite.

An acid may have no effect, the mineral being insoluble, as quartz.

In dissolving (or combining) minerals with HCl you change them to chlorides; with HNO_3 to nitrates; with H_2SO_4 to sulphates.

Carbonates will effervesce with HCl (and other stronger acids) some with the cold acid, others only with the hot acid.

CHAPTER XXVI.

BLOW-PIPE ANALYSIS.

The analysis of many m'neral compounds requires a greater heat than that of an alcohol lamp. This is obtained by means of the *blow-pipe*, a simple instrument by which a strong current of air is added to the flame of a lamp or gas-jet, increasing its heat by a freer supply of oxygen, and at the same time directing and concentrating the flame upon the specimen.

A simple curved brass blow-pipe, costing 20 cents, answers the purpose. A stearine candle, with a thick wick, furnishes a good flame; but a properly arranged Bunsen burner is better. Fig. 28½.

Charcoal is generally used as a support for the substances to be heated. It is best to buy the coal which is specially prepared for this purpose.

Platinum wire and foil must be on hand. Platinum makes excellent instruments for holding specimens in the blow-pipe flame, because it does not melt.

The wire, No. 27, is cut into pieces about two inches long. These should be fitted with wooden or glass handles. Hold the end of a small glass tube, two inches long, in the Bunsen flame, turning it as it grows hot. As the glass softens, the opening in the tube grows smaller; when just large enough to admit the

platinum wire, insert about a quarter of an inch of it into the tube, and finish melting it together.

Bend the free end of the wire into a round loop, an eighth of an inch in diameter, and you have a service

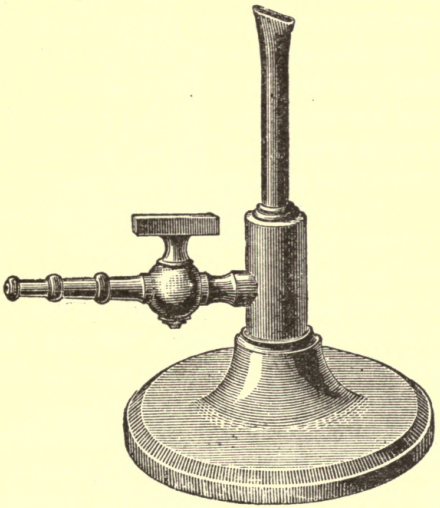

Fig. 28½.

able little instrument. A forceps with platinum points is so convenient, as well to repay its cost.

Care must be taken not to use any platinum instruments for testing lead, tin, zinc, or other easily fusible ores, for they make an alloy with platinum, and render it easily destructible.

Lay a sheet of paper under the lamp or burner before beginning work. It will catch fragments dropped dur-

ing the examination. Record all observations in a note-book.

You must learn to blow a continuous blast through the blow-pipe, while breathing naturally. This is accomplished by inflating the cheeks and allowing them to contract, expelling the air while you are drawing in your breath. It is difficult at first, and you should get some one to show you how to do it if possible.

Hold the tip of the blow-pipe in the middle of the flame, just above the wick, if you use an alcohol lamp or a candle; do not push it quite so far into the flame of a Bunsen burner. Blow gently, so as to cause the flame to taper away from the blow-pipe to a fine point. The brightness of the flame should disappear, and you should observe a fine blue cone surrounded by a very faint flame.

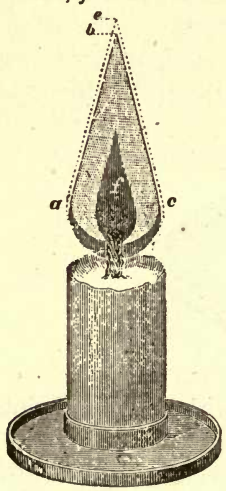

Fig. 29.

When you can produce this flame and maintain it steadily for several minutes, blowing from the cheeks and not from the chest, you have mastered the art of blowing. Fig. 29 represents a candle flame, *f* is the dark central portion filled with gases formed from the heated stearine; *a b c* is the brilliant flame where the combustion is incomplete, and the particles of carbon are incandescent; *a e c* is the faint outer flame caused by the

complete combustion of the carbon with hydrogen and oxygen. Upon introducing the blow-pipe and blowing gently, the flame takes the position shown in Fig. 30.

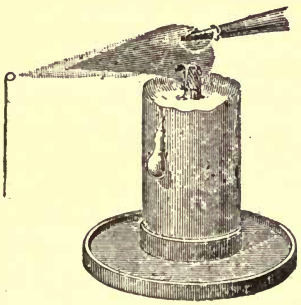

Fig. 30.

The increased supply of oxygen causes more perfect combustion, as shown by the disappearance of the brilliant flame, and the formation of the blue cone, surrounded by a still fainter conical flame which extends beyond it.

Substances held just beyond the blue cone in this faint outer flame are heated in contact with the oxygen of the air, and many metals thus become oxidized. This outer flame is therefore called the oxidizing flame.

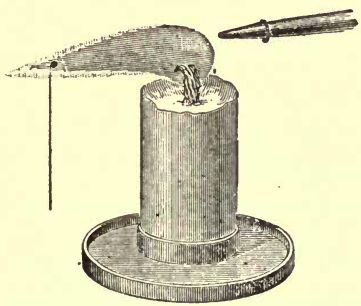

Fig. 31.

If the tip of the blow-pipe is held just outside the flame, and the flame gently directed outward and downward, Fig. 31 the flame keeps its yellow colors, and if an oxide of a metal is held in the end of this flame it becomes partly or entirely deoxidized or *reduced.*

This then is the *reducing flame.*

O. F. stands for "oxidizing flame." This flame is used not only for changing metals or their compounds into oxides, but also to test the fusibility of minerals. The hottest part is the tip of the blue cone.

R. F. stands for "reducing flame." It is chiefly used to separate the metals from their compounds.

The Bunsen flame should be about an inch and a half high for the O. F., and half an inch higher for the R. F.

FIRST EXPERIMENTS.

Direct the O. F. upon a splinter of cryolite.

Observe how readily it melts. Try a small splinter of celestite. It fuses, though less readily, and shows its characteristic crimson flame. Place a drop of water on red litmus-paper, and dip into it the fused splinter of celestite; notice that the litmus turns blue. What does this indicate?

Now try to melt a fragment of quartz. If you can melt it in the slightest degree you perform a feat hitherto deemed impossible.

Try next a bit of galenite. In the platinum forceps? Not unless you wish to ruin the instrument!

Make a slight excavation in a piece of charcoal about half an inch from the end; in this place the galenite and direct the tip of the O. F. upon it. Stop occasionally to observe the coating formed on the coal, and to smell the fumes that arise. Notice that the color of the coating near the specimen quickly changes from orange to light yellow as you remove the coal from the flame, and that it grows still lighter as it gets cold. Galenite, you re-

member is lead sulphide. Some of the sulphur has *volatilized*, combining with oxygen to form the odorous fumes. Some of the lead has united with oxygen, forming the yellow oxide deposited near the specimen or *assay*; another part of the lead has been changed into lead sulphate, and is deposited farther away on the cooler portion of the coal. Upon the remainder of the assay direct now the R. F., until a bright little drop, or *globule* is formed. Let this drop on a piece of asbestus paper, or let it cool on the coal.

It is no longer brittle galenite, but malleable lead, as you can prove by a blow of a hammer.

These simple experiments have shown the use of the blow-pipe:

To test the fusibility of minerals;

To exhibit colors imparted to the flame;

To oxidize and to reduce metals;

To cause the formation of coating; or *sublimates*, on charcoal—often an important aid in determining a mineral.

DEGREES OF FUSIBILITY.

We may call *very easily fusible.*—Minerals whose fragments melt in the alcohol or gas flame without a blow-pipe; as stibnite, and halite. Fusibility 1.

Easily fusible.—Minerals that melt in the blow-pipe flame in fragments as large as a grain of wheat; as pyrites. Fusibility 2.

Fusible.—Minerals melting less readily, but still visibly affected in fragments as large as wheat grains; as celestite. Fusibility 3.

Fusible with difficulty.—Minerals that fuse only in very small splinters; as most varieties of hornblende. Fusibility 4.

Almost infusible.—Minerals that are rounded only on the sharp points, and edges of minute splinters, after long and strong heating, as orthoclase. Fusibility 5.

Infusible.—Minerals not affected by flame; as quartz. Fusibility 6.

All the minerals named above are included in the collection accompanying this book, and you should become familiar with their action before the blow-pipe before attempting others.

Other phenomena than fusing are sometimes observed when blow-piping, *e. g.* swelling, branching, decrepitating, brightly glowing (especially lime, strontia, zinc, and tin). Careful notes should be made of everything observed.

EXAMINATION ON CHARCOAL.

Easily fusible metals must be examined on charcoal, as we saw in the case of galenite. If the specimen decrepitates so strongly as to fly to pieces when touched by the flame, powder it. Moisten with a drop of water as much of the powder as you can take on the point of your knife-blade, mix to a paste, and place the mass in a little excavation made near the end of the charcoal. In directing the flame upon the assay blow gently until the moisture is expelled.

Try stibnite. It fuses readily, imparting a greenish hue to the flame, and yielding dense white fumes which form a white coating on the coal,

Touch this coating with the tip of the R. F. and it disappears, again tinging the flame with green. This proves that the sublimate of antimony oxide is volatile.

Try sphalerite. The fragment does not fuse, though it may form a yellow coating near the assay, turning white as it cools.

You remember that when limonite was heated in a closed tube a large quantity of water was driven off, and the mineral afterward had the appearance of hematite. Repeat the experiment with the blow-pipe. The same result is produced more quickly under the intenser heat. Do not stop this time when the limonite has been reduced to hematite; continue the heating, and observe that the assay grows still darker, and at last partially fuses. Test it now with a magnet. Unlike both limonite and hematite, it is attracted.

Yet it is not pure iron, for a smart blow of the hammer crumbles it to powder. It is reduced to that compound of iron called *magnetite*, specimen 12. It is believed that this series of iron ores has been formed, in nature, on the same principle which has been illustrated in our experiments; *i. e.*, limonite, deposited in the earth from a solution of iron, has afterward been converted into hematite by heat, and hematite has in the same way been reduced to magnetite.

Test now a few grains of iron-rust, which you can scrape from an old stove-pipe, or other iron. Does its streak resemble limonite? It is practically the same mineral. What are the results of heating it? Test the assay, after heating, with a magnet.

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Repeat the experiment of burning iron in oxygen. Test with the blow-pipe a fragment, or some powder, of the resulting black oxide of iron. Test it with the magnet before and after heating.

REAGENTS.

We have frequently used litmus-paper in order to detect an acid or an alkaline reaction; a change of color from blue to red indicating the presence of an acid; while minerals containing as their base calcium, strontium, barium, or magnesium all have the same alkaline reaction after being strongly heated which we observed in the case of celestite.

Cobalt is a reddish-white metallic element, of no special utility, except for the brilliant colors that are obtained from its various compounds. To glass these compounds impart a magnificent tint which is easily recognized as cobalt blue. With other substances than glass other colors are produced, so that by the color obtained by moistening different minerals with a solution of some cobalt compound and then heating them it is often possible to determine their nature and their names.

The most convenient preparation of cobalt for this purpose is a solution of one dram of the nitrate, $Co(NO_3)_2$, in an ounce of water.

Heat another fragment of sphalerite on charcoal until you obtain the yellow coating or sublimate on the coal next to the assay. Put a drop of the cobalt solution on the sublimate, and also upon the assay, if that has changed to a light color through heating. Direct

the O. F. again upon the assay; when cool, the sublimate or part of it, and probably also the assay, will be bright green.

If you fail to get any coating or sublimate by simply heating the sphalerite, pulverize a fragment in the mortar, mix the powder in the palm of your hand with an equal quantity of pure sodium carbonate,—or of ordinary baking soda if free from sulphur—moisten the mixture and expose it to the O. F. It will melt and run together, but after a time will spread out again and sink into the coal. Keep playing upon it with the hot point of the flame until you have a distinct coating beyond the assay; then treat with the cobalt solution as before. Repeat this test with all the ores of zinc you may have. You should keep a small bottle of this cobalt solution on hand. Always make the test in the same way as you have with sphalerite, *i. e.*, heat a fragment or some powder of the mineral strongly on charcoal, in the O. F; moisten it or its sublimate with a drop of the cobalt solution, and then heat again.

Zinc oxide then becomes yellowish-green.

Tin oxide becomes bluish-green.

Magnesium oxide, magnesia, becomes pink.

Aluminium oxide, alumina, becomes blue. Try kaolinite.

Silicon oxide, silica, becomes faint blue. Try quartz.

Strontium oxide, strontia, becomes gray. Try celestite.

Calcium oxide, lime, becomes gray. Try chalk.

You will be interested to note in this connection a

few characteristic coatings which the O. F. will produce on charcoal without the assistance of any reagent. White, very volatile, gray fumes, indicate arsenic white near the assay, blue farther off, dense white fumes indicate antimony.

Purplish-brown, indicates silver. In lead rich in silver the color may be brighter red, or the brown may appear in spots and streaks amid the yellow oxide of lead.

Lemon-yellow while hot, sulphur yellow when cold, white and bluish-white at a distance, volatile, indicates lead.

Always wait until the coating is completely cool before judging its color.

Do not mistake the white ash of the charcoal for a sublimate. Sometimes it is necessary to fuse the assay with soda in order to get the sublimate.

Do not fail to notice the odor given off during heating.

ROASTING.

Substances containing sulphides or arsenides must be *roasted* before they can be treated with soda or any other flux. Grind small fragments of iron pyrites to fine powder. Moisten the powder and transfer to a shallow cavity in the charcoal. Heat very gently at first to prevent the powder from melting or flying away. After some of the sulphur has been driven off use a stronger blast—the O. F. and the R. F. alternately.

Continue as long as there is the slightest smell of sulphur, no matter how long it may take. Turn the assay on the other side and treat as before. Again pulverize the assay in a mortar. The powder is no longer greenish, but red like that of hematite. It is indeed, now, an oxide of iron. As some sulphur probably remains, roast it again on both sides. Test the powder with a magnet and carefully preserve it for the following experiments.

TESTING WITH FLUXES.

Heat your platinum loop and dip it into powdered borax. Hold it over the flame of the lamp or gas and watch the melting of the borax after "intumescence;" but to prevent particles of the borax from flying away, hold it *in* the flame.

Add more borax until the loop is filled. Then direct the O. F. upon it with the blow-pipe, until it becomes a transparent glass, free from bubbles. This is a borax "bead." Place a very little of your roasted assay upon the bead. If it will not stay, smoke the bead á little over the flame, or heat it again with the blow-pipe, and flatten it gently between the forceps or on an anvil. Heat the bead over the lamp-flame; just enough to cause the powder to stick to it. Now hold the loop in the O. F., beyond the point of the blue flame, as shown in Fig. 30. The powder begins to whirl about in the melting glass, but it, too, soon dissolves, and the glass becomes clear and smooth.

Hold it between your eye and a window; it is yellow,

but becomes colorless as it cools. Add a little more of the powder. Dissolve as before in the O. F. This time it appears red while hot, and yellow when cold. Heat it now in the R. F., holding it farther in the flame than shown in Fig. 31, as your loop is larger. The bead assumes a greenish tint. Add still more powder. The bead becomes brown or reddish yellow in the O. F., and in the R. F. should assume the peculiar bottle-green of junk-bottles.

A bead which contains as much of any assay as it can dissolve is said to be *saturated.* After a bead is saturated, you may again dilute it. Melt the bead and fling it into a clean cup or mortar, by striking the hand which holds the platinum loop, on the table. Part of the assay will still adhere to the loop; add to thi. enough borax to make a new bead. This will resemble the second one you made. Strike this off, too; dilute again with borax, and you will have again a yellow bead that becomes colorless as it cools.

Place on charcoal a piece of tinfoil, and on this lay the two beads that were shaken off. Expose the beads and the foil under them to the R. F.; the bead should become decidedly green, although in order to see the color you may have to pinch the bead, which clings to the coal, upward with the forceps so that you can look through it.

*Care of the platinum wires.** The test being finished, remove the bead by gentle taps with a hammer, remove all clinging particles, and keep the wire in a bottle con-

*Keep several platinum loops on hand.

taining enough diluted sulphuric acid to cover the loop. Before using again rinse in water. If the new bead of borax is not perfectly clear and colorless both when cold and after heating, throw it off and make a new one. If after several trials the color still persists, make a bead with soda, shake it off, and put the loop into the acid bottle for a day. Should you find, while testing a mineral in the borax bead, that a metallic globule is formed, shake the bead off immediately and transfer it to charcoal.

Test other specimens of iron ores in the borax bead. The oxides and carbonate need no roasting.

Try a little powdered chalk in the bead. It effervesces as it dissolves and forms a colorless glass. Add a small fragment instead of the powder; the effect is the same. Keep on adding and dissolving, examining the bead each time. When nearly saturated it becomes translucent instead of transparent on cooling, and assumes a crystalline structure. Watch it with a magnifying glass, for the process is extremely interesting. Add a little more chalk and the bead becomes at last almost pure white, and opaque. If you direct the flame upon the bead in little puffs or flashes just before the point of saturation is reached it will become opaque even in the flame. This is called *flaming* the bead.

Throw off part of the assay, dilute the rest with a little borax so that it becomes clear again and try the flaming. It is chiefly the alkaline earths that become thus opaque by flaming.

Compare the action of kaolin with that of chalk.

Make a borax bead with finely powdered quartz. It, too, dissolves to a clear glass, but more slowly.

From these experiments we learn that borax causes even the most refractory metals to melt in the heat of the blow-pipe.

Substances which have this property are called *fluxes*.

The following are the principal reactions with the borax bead:

Colorless.—The alkaline earths, aluminum, silica; in O. F., and R. F.

Yellow when hot, colorless when cold—in O. F., zinc, cadmium, lead, bismuth. In R. F., these substances become gray from the reduced metal, but on further heating grow clear again.

Milk-white—silver, when the bead is saturated, in O. F.; gray, then clear in R. F.

Yellow in O. F.—iron, titanium, tungsten, molybdenum, chromium (greenish yellow).

Red to brown in O. F.—hot; iron, uranium, chromium:—cold; nickel, manganese.

Red in R. F.,—copper, when the bead is saturated and cold; colorless when hot.

Violet in O. F.,—nickel, hot (reddish-violet); manganese, (deep violet).

Blue—cobalt, O. F., and R. F.; copper, O. F., cold, greenish-blue.

Green—copper, O. F., hot; R. F., bottle-green; chromium, vanadium, O. F. and R. F., cold; uranium, R. F., yellowish-green.

EXAMINATION WITH SODA.

Soda beads in the platinum wire are not often used, but it is worth while to make a few. The bead itself is interesting. After it has become well fused, notice how it begins to effervesce, as you pass it from the O. F. to the R. F., also that it is transparent while hot, and becomes translucent when cool.

Add some quartz powder; it easily dissolves, with effervescence, and the bead on withdrawal from the flame loses its transparency. Continue adding and dissolving quartz, and presently you reach a point where the bead remains transparent even after cooling.

Try orthoclase, and you should secure the same result.

From this we understand that silica dissolves readily and with effervescence in soda, and that, if the proportion of soda is not too great, the resulting glass will be clear. Indeed, our window-panes and common glassware are for the most part silicate of soda, or of soda and lime.

In most cases the treatment of mineral compounds with soda is conducted on charcoal.

Mix some powdered quartz with about three times its bulk of soda and transfer the paste to the coal. Treat it with O. F. for a considerable time, trying to collect the mass into a globule. You will hardly succeed. Probably the assay will spread out more and more, and partly sink into the coal. Now make another mixture, taking *equal* parts of quartz and soda. You should succeed without much trouble in forming this mixture

into a bead. Continue playing the O. F. upon it, causing it to revolve until you get rid of nearly all bubbles. You should get a transparent, nearly colorless bead.

Next try a small splinter of orthoclase. Place it in a shallow depression in the charcoal and put a little soda around it. Blow gently, as you always should at the beginning. See the soda fuse, and begin to attack the mineral; effervescence marks the beginning of the fusion; the edges and corners are first dissolved. Add a little more soda and heat again, the assay spreads out to take in every grain of the salt, and then rounds up to form a globule, with the inner part, probably, yet undissolved. Other additions diminish the undissolved part. and the bead may assume a dark appearance from some impurities taken up. Add much soda and the mass will spread out. It takes longer to do this work than to describe it, but the experience gained during the faithful and patient performance of the experiment will be worth all it costs.

You should still have some of the powder from the roasted pyrites. Mix all that is left with about three times its bulk of soda, moisten, and transfer to the charcoal. Give it a long and strong blast with the R. F., until all the *life* has gone out of the mixture, *i. e.* as long as you can observe any effect from the heat, and even longer.

Cut out the black mass with some of the coal underneath it, and grind it to a powder in a mortar, adding a little water. Set the mortar with its contents in a cup or basin, and pour water in the vessel until it covers the

mortar, and a little more. The lighter portion of the assay will now rise from the mortar, and float on the water; and by moving the mortar about under the water all but the heaviest sediment can be floated out.

Now remove the mortar, pour the water from it, and carefully transfer the sediment to a piece of paper. When dry, examine the little particles with a magnifying-glass. Test them with a magnet. It is a good plan, by the way, to fit the ends of your magnet with "stockings" of tissue paper reaching half way up. Magnetic substances will be attracted through the paper, and they will drop off when you pull the magnet partly out of the stocking. You have now reduced the oxide to metallic iron.

Powder some galenite, mix with plenty of soda, and treat with the R. F. Let the globules run together, and observe the action of the large globule of lead as it solidifies. When it grows cool take it off and hammer it. Observe how much more lead you get in this way from a certain quantity of galenite than by reducing the galena without soda.

The last two experiments illustrate the use of soda as a flux for reducing metals from their compounds. If any metal does not form a globule, but behaves like iron, forming a *slag*, this slag must be treated as the iron was, that is, powdered and separated by the floating process.

If then the metal cannot be recognized from the sediment in the mortar, this must be dissolved in a borax bead, that its nature may be revealed by the color imparted to the bead.

We must remember, too, that the treatment with soda on charcoal may be used for getting a coating or sublimate which may aid in determining the metal.

All reductions on coal must be conducted with the R. F.

If you suspect that the dark mass or *hepar* resulting from heating a metal on coal with soda contains sulphur, crush it and place a fragment on a bright silver coin, adding a drop of water. If sulphur is present it will unite with the silver forming silver sulphide, and leaving a brown stain on the coin. Be sure, however, that neither your soda nor your fuel contains sulphur.

Filtering and decanting. To separate solutions from an undissolved sediment, or from a precipitate, they are passed through filter-paper placed in a funnel; or, if the liquid stands clear above the sediment, may be simply poured off, or *decanted.* If the sediment or precipitate is to be further tested it should be washed either by pouring water over it as it lies on the filter-paper, and letting it pass through until all acid or alkali is removed, or by repeated shaking up and decanting with water.

CHAPTER XVII.

SUGGESTIONS FOR FURTHER STUDY.

No experiments have been described in this book which cannot be successfully performed without assistance by every student, no matter whether he has had any previous training or not, if he will undertake them resolutely and patiently, and *in the order given.*

Those who have worked their way with us to this point will wish to continue their study of this fascinating science, until they attain the power of recognizing at sight, or of determining by analysis, every specimen they may meet. A description of twenty-four more common minerals is therefore added for practice. It must be borne in mind, however, that the determination of the names of minerals is by no means the chief aim of the earnest student. It is rather to make himself thoroughly acquainted with the properties of minerals, and to develop and strengthen his powers of observation and reasoning. Carefully to observe, and accurately to describe your specimens, should therefore be your main object.

It is almost wholly because the processes required to determine minerals necessarily enforce this patient and exact personal observation, that they become a valuable means of education and training.

You will do well, now, to procure the specimens whose descriptions follow.

1. *Native copper.* H.3; lustre and streak, metallic; color red; occurs in masses, in crystals, and in beautifully branched or *arborescent* forms, sectile, malleable, very tough, heavy.

Comp.—Cu; B.B. fuses to a globule, try tests for Cu.

2. *Cinnabar.* H.2-2.5; lustre, adamantine to dull; streak, scarlet; color, red to dark brown; in crystals, or massive; brittle; very heavy; fracture uneven.

Comp. HgS; B. B. volatilizes; Hg and S reactions in c.t. with and without soda. Cinnabar is the principal ore of mercury, which is also found native.

3. *Millerite*, (*nickel pyrites*). H.3-3.5; lustre, metallic; color, bronze yellow; streak, gray; in clusters of fine needle-shaped crystals, or in fibrous crusts; brittle; rather heavy.

Comp.—NiS; B. B. on charcoal fuses easily to a brittle, magnetic globule; with borax bead gives Ni reaction; roasted and fused with soda on charcoal yields a brittle globule of nickel.

4. *Pyrrhotite.*—H. about 4; lustre, metallic; color, bronze yellow, tarnishing to copper red; streak, gray, generally massive; brittle; magnetic.

Comp.—Fe_7S_8; B. B. on charcoal fuses to a magnetic globule; yields little sulphur in c. t.; softer than iron pyrites and harder than copper pyrites; *magnetic.*

5. *Chalcopyrite* (*copper pyrites*). H. 3.5; lustre, metallic; color, brass yellow, often tarnished to deep yellow, red, blue, etc.; streak, greenish black; crystal-

lized or massive; fracture, uneven; brittle; *easily scratched with a knife.*

Comp.—$CuFeS_2$; B.B. on charcoal fuses easily to a magnetic globule; fused with soda on charcoal yields a globule of copper and iron. Moisten the globule with HCl., and examine the flame color at the point of the O. F. Crush the globule to powder and dissolve in a few drops of nitric acid; both copper and iron will be dissolved. Add ammonia until the iron is precipitated as a reddish brown powder; the liquid will then be blue. Also roast some of the ore, or make another globule of it with soda. Pulverize, and saturate a borax bead with the powder; examine the bead and throw it off; add more borax to what remains on the wire, examine it both hot and cold, and throw off again; repeat this until the bead becomes almost colorless, and make a record of your observations.

6. *Cerargyrite* (*horn silver*).—H. 1.5; lustre, resinous; streak, shining; color, pearl-gray, brownish when weathered; in very small cubical crystals or massive; fracture uneven; easily cut with a knife.

Comp.—AgCl; melts in the candle-flame. B. B. on charcoal with O. F. yields little globules of silver. Better, however, fuse on charcoal with soda in R. F., and collect all the silver into one globule. Observe the coating on the coal, which may show, besides the characteristic purple brown of silver, a yellow sublimate, indicating the presence of lead.

7. *Zincite.*—H. 4–4.5; lustre, brilliant; streak, orange yellow; color, bright red; generally in granu-

lated or foliated masses; cleavage, perfect; brittle; rather heavy.

Comp.—ZnO. B. B. infusible; turns black, but after cooling turns red again. On charcoal, with and without soda, gives zinc coating.

8. *Menaccanite* (*titanic iron*).—H. 5–6; lustre, metallic or sub-metallic; color, black; streak, black; in grains, plates, sometimes large tabular crystals; cleavage, distinct; rather heavy; *magnetic.*

Comp.—$(Ti\ Fe)_2O_3$; B. B. infusible; fused with soda on charcoal, and dissolved in HCl, gives Ti reaction.

9. *Franklinite.*—H. about 6; lustre, metallic; color, iron black; streak, reddish brown to black; massive or in crystals; cleavage, indistinct; brittle; fragments slightly magnetic.

Comp.—$(FeZnMn)_2O_4$; B. B. infusible; with soda on charcoal, white zinc coating, turning green if treated with cobalt solution. In soda bead assumes a green color in O. F. Both reactions are more distinct if a little borax is added to the soda.

10. *Chromite.*—H. 5–5; lustre, metallic, or sub-metallic; streak, brown; color, iron black; generally massive; breaks with even but rough surfaces; not magnetic, unless in very small particles.

Comp.—$FeCr_2O_4$; B. B. infusible; with borax bead Cr reaction.

11. *Cassiterite.*—H. 6–7; lustre, splendent, sub-metallic, dull on exposed surfaces; streak, brownish to gray; color commonly brown to black; in crystals; of-

ten bean-shaped, fibrous, also massive and compact; fracture uneven; brittle; heavy.

Comp.—SnO_2; B. B. infusible, with soda on charcoal yields tiny globules of tin. *See Sn, p.*—The principal ore used in the manufacture of tin.

12. *Pyrolusite (black oxide or dioxide of manganese)*—H. 1–2. 5; lustre, metallic to earthy; streak and color; black; soils the fingers; generally massive, columnar, radiated; brittle, rather heavy.

Comp.—MnO_2; B. B. infusible; very little imparts an amethyst color to the borax bead. Dissolve in HCl, heating the solution slightly, and observe the odor of chlorine.

13. *Beauxite.*—H. about 3; lustre, earthy; color, white, red to brown; streak, white to red; composed of small round concretions like little peas, (pisolitic), imbedded in a compact matrix; fracture, uneven, brittle.

Comp.— $Al_2Fe_2\ O_3+2H_2O$; B. B. infusible; decrepitates strongly; in c. t., yields water. Try reaction for alumina; for iron.

14. *Spodumene.*—H. 6.5-7; lustre, pearly; color, grayish to greenish white; streak, white; massive, also often in large crystals; cleavage perfect; brittle.

Comp.—$Li_6Al_8Si_{15}O_{45}$; B. B. fuses easily with intumescence, first to a white mass, then to a clear glass, showing, at least for a few moments during the treatment with O. F., the lithia reaction. Try the tests for alumina and silica.

15. *Willemite.*—H. 5.5; lustre vitreous to resinous, color, brown, red, yellow, green, also colorless; streak,

white; translucent, seldom transparent; generally compact, fine grained, sometimes in small crystals; fracture uneven.

Comp.—Zn_2SiO_4; B. B. fuses with difficulty and turns white; no water in c. t.; with HCl gelatinizes perfectly; an important ore of zinc.

16. *Cinnamon garnet*, (massive).—H. about 7; lustre, vitreous to resinous; color, commonly red or brown, but also of almost every other color; generally in crystals; also massive; brittle; fracture uneven.

Comp.—$Ca_3Al_2Si_3O_{12}$; B. B. fuses less easily than vesuvianite (the next specimen), and without swelling, to a dark glass.

17. *Vesuvianite* (*idocrase*). H. 6.5; lustre, vitreous to resinous; color, brown to green; streak, white; subtransparent to nearly opaque; in prismatic crystals, also massive; fracture uneven, brittle.

Comp.—$(CaMg)_8(AlFe)_4Si_7O_{28}$; B. B. fuses with swelling to a dark greenish or brownish glass, which shows alkaline reaction. Fuse a few fragments thoroughly on charcoal and dissolve in HCl; boil it well down and notice the gelatinizing of the mineral.

18. *Scapolite*, (*wernerite*). H. 5-6; lustre, vitreous to pearly; color, usually light gray, blue, lilac, green, or pink; streak, white; transparent to nearly opaque; crystallized or massive; often with a fibrous or columnar appearance; fracture, inclining to conchoidal; brittle.

Comp.—$CaAl_2Si_2O_8$; B. B. fuses rather easily, and with intumescence to a whitish, bubbly glass, which cannot be further fused; alkaline reaction.

19. *Elæolite* (*variety of nephelite*, H. nearly 6; lustre, vitreous to greasy; color, grayish; streak, white; translucent; massive or in large crystals; fibrous appearance on broken surfaces; brittle.

Comp.—$(NaK)_2Al_2Si_2O_8$ (Dana) B. B. fuses quietly, but not easily, to a white glass resembling that of scapolite. Gelatinizes with HCl without previous fusion. No alkaline reaction.

20. *Sphene* (*titanite*). H. 5-5.5; lustre, vitreous to resinous. color, brown to black, sometimes yellow or green; streak whitish; transparent to opaque; generally in oblique, thin-edged crystals, sometimes massive; cleavage often perfect in one direction, giving the crystals a laminated appearance; brittle.

Comp.—$CaTiSiO_5$; fuses with intumescence to a blackish mass. Try test for Ti.

21. *Calamine*.,—H. 4.5–5; lustre, vitreous to pearly; color, various, commonly light brown; streak, whitish; transparent to translucent; massive, or forming botryoidal, stalactitic, or radiated incrustations lined with crystals (generally in fissures of rock); cleavage, perfect in crystalline varieties; fracture, uneven to rough in granular or earthy varieties; brittle.

Comp.—$Zn_2Si\ O_4 + H_2O$; B. B. fuses with difficulty, turning white. In c. t. yields water. With soda on charcoal gives zinc coating; treated with cobalt solution green, assay blue. Dissolves in HCl, leaving white residue. Gelatinizes on boiling down the acid. An important ore of zinc.

22. *Strontianite*.—H. 3. 5–4; lustre, vitreous; color,

white to yellowish and greenish; streak, white; generally massive, fibrous in parallel or radiated lines, also globular; transparent to translucent; fracture, uneven; brittle.

Comp.—$SrCO_3$; B. B. fuses with difficulty, white sprouts appearing on the points and edges; colors the flame crimson; heated, shows alkaline reaction. Try reaction for carbon dioxide. Strontianite is used in the manufacture of fireworks.

23. *Malachite* (*green copper carbonate*).—H. 3.5–4; lustre, silky to earthy; streak, pale green; color, green; generally opaque, sometimes translucent; massive or forming incrustations which may be mere films, or thicker botryoidal, stalactitic, or spongy masses of compact and fibrous structure; fracture, uneven; brittle.

Comp.—$Cu_2CO_4 + H_2O$; B. B. easily fusible to a black globule; observe the green flame beyond the globule. Moisten with HCl and notice the flame. On charcoal fuses to a globule of copper; if there is difficulty, add a little soda, running it into a bead with the O. F., and then reducing it in the R. F. Try other reactions for copper.

24. *Azurite* (*blue copper carbonate*).—H. 4; lustre, vitreous or dull; color and streak, blue; translucent to opaque; in crystals, or massive, also compact, earthy; brittle; much like malachite except in color.

Comp.—$Cu_3C_2O_7 + H_2O$; behavior with reagents like malachite. Both minerals are valuable ores of copper. Malachite is often worked into mantels, tabletops, and various ornaments.

REFERENCE LIST

of the principal elements in minerals and their characteristic reactions.

Ag.—*Silver.* B. B. Heated on charcoal in O. F. forms a reddish brown coating. Dissolved in nitric acid forms a whitish cloud (precipitate of Ag Cl) which settles and turns dark on exposure to light, and which may be dissolved again by the addition of ammonia. Most silver ores heated on charcoal, alone or with soda, yield a silver globule.

Al.—*Aluminium.* (Al_2O_3, alumina) B. B. A fragment strongly heated on charcoal, or in the forceps, then moistened with cobalt solution and heated again, assumes a beautiful blue color.—*All color reactions should be examined by daylight.*—Sometimes it is necessary to pulverize before heating. In dark-colored specimens the blue does not always appear distinctly; if these are soluble in acids, and ammonia be added to the acid solution, the alumina will be precipitated, and after drying can be tested with the cobalt solution.

As—*Arsenic.* B. B. garlic odor, gray fumes, gray coating on charcoal, distant from the assay and easily driven away by the flame. In o. t. yields a white crystalline sublimate; in c. t. a black brilliant sublimate.

Au—*Gold.* Imparts no characteristic color to flame or flux. If native is easily recognized by its color, malleabilty, and insolubility in any one of the acids.

Ba—*Barium.* (BaO, baryta) B. B. yellowish-green flame, shown by all minerals containing barium except silicates. If soluble in HCl, an addition of H_2SO_4 causes a heavy white precipitate. $BaSO_4$.

Bi—*Bismuth.* B. B. on charcoal alone or with soda, gives a characteristic orange-yellow sublimate. Treated with equal parts of potassium iodide and sulphur, a beautiful red precipitate of bismuth iodide is formed.

Bo—*Boron.* B. B. Intense yellowish-green flame, especially if moistened with H_2SO_4.

C—*Carbon.* (CO_2, carbonic acid, or anhydride) Is set free, when present in minerals, by HCl and all the stronger acids causing effervescence. Sometimes the mineral must be pulverized, and sometimes the solution must be heated.

Ca—*Calcium.* (CaO, lime), B. B. yellowish-red flame. Many minerals containing lime give an alkaline reaction after being highly heated.

Cl—*Chlorine.* B. B. Prepare a bead with the salt of phosphorus, (a phosphate of sodium and ammonium), saturate it with copper oxide, and add some of the powdered mineral; if Cl is present the blow-pipe flame will be colored for a moment intensely blue. (Try halite, specimen No. 28).

Co—*Cobalt.* B. B. imparts to the borax bead a beautiful blue color. Try also with cobalt solution.

Cr—*Chromium.* B. B. imparts an emerald-green color to the borax bead, (cold) in both O. F. and R. F.

Cu—*Copper.* B. B. green flame; blue if moistened with HCl, borax bead, O. F., green, R. F., red, if highly saturated. Add a grain of salt (NaCl), and observe the flame. Most copper ores are reduced to metallic copper when fused with soda on charcoal.

F—*Fluorine.* If the powdered mineral be laid on

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glass, moistened with one or two drops of H_2SO_4, and set aside for a day, the glass will be *etched.* This corroding action will be more pronounced if the powdered mineral containing F is mixed with bisulphate of potassium, and heated in the c. t.; a strip of Brazil-wood paper inserted in the mouth of the tube will turn straw-color.

Fe—*Iron.* B. B. besides the tests already given on p. 204, try the following: Powder some of the mineral, add dilute HCl, and heat; let the sediment settle, and decant the clear solution; dilute this with more water; add a drop of a solution of yellow prussiate of potash If iron is present, the solution will soon turn blue. (Prussian blue.)

Hg—*Mercury.* Its compounds yield a metallic coating when heated with soda in the c. t. The globules of mercury are sometimes so small that the coating appears white. Use the magnifying-glass, or brush the tiny globules together with a strip of paper. Sulphide of mercury yields a black, volatile sublimate, red when rubbed.

K—*Potassium.* B. B. is often recognized by the violet color it imparts to the flame. If sodium is present this color does not appear.

Li—*Lithium.* B. B. purplish to carmine flame, as in lepidolite.

Mg—*Magnesium.* (Mg O, magnesia). B. B. generally indicated by the pink color which its compounds assume when highly heated, and treated with cobalt solution.

Mn—*Manganese.* B. B. imparts a beautiful amethyst color to the borax bead, and assumes a bright green color when fused with soda.

Na—*Sodium.* B. B. colors the flame intensely reddish yellow, and seems to make it larger.

Ni—*Nickel.* B. B. in O. F. colors the borax bead violet when hot, reddish brown on cooling; in R. F. colorless with metallic specks. Nickel ores are reduced to a magnetic mass on charcoal.

P—*Phosphorus.* B. B. characteristic green flame, especially when moistened with H_2SO_4. A pulverized phosphate fused in the c. t. with a bit of metallic magnesium or sodium, evolves the disagreeable odor of phosphoretted hydrogen. (Try apatite.)

Pb—*Lead.* B. B. when heated with soda on charcoal compounds of lead yield a metallic lead globule; a yellow volatile coating appears on the coal, deeper near the assay, and shading to white and bluish.

S—*Sulphur.* B. B. on charcoal most sulphides are recognized by their odor; bisulphides, as FeS_2, by the sublimate of S in the tube. Sulphides dissolve in HCl with effervescence yielding fumes of sulphuretted hydrogen—(rotten-egg odor). For the "hepar" test by which the presence of S in sulphates can be detected, see p. —.

Sb—*Antimony.* B. B. on charcoal, generally recognized by its dense white fumes and white sublimate. Same in o. t. Bluish flame.

Se—Selenium. B. B. steel gray, volatile sublimate; disagreeable horse-radish odor.

Si—*Silicon.* [SiO_2, Silica] B. B. heated with soda in platinum wire, or on charcoal, fuses with effervescence to a clear glass; if too much soda is used a *slaggy* mass results.

Sn—*Tin.* B. B. with soda on charcoal yields globules of metallic tin, often so minute as to be detected only with a lens. By crushing the fused mass with the adhering coal between paper and heating again, adding a little soda, the globules can generally be collected into a larger globule. The operation requires skill and care. Tin yields a faint yellow sublimate,—not volatile,—which becomes bluish green when treated with the cobalt solution.

Sr—*Strontium.* B. B. crimson flame, especially if moistened with HCl.

Ti—*Titanium.* B. B. colors a bead of the salt of phosphorus violet, unless too much iron is present. In that case wrap the bead in a small piece of tinfoil, transfer to charcoal, and apply the R. F. The iron then will unite with the tin in a metallic globule, leaving the Ti in the bead, when the violet color will appear.

Tu—*Tungsten.* Tungstic acid colors the salt of phosphorus bead blue in R. F. If iron is present the bead is brownish red; melting on charcoal with tinfoil or zinc the bead is turned blue.

Zn—*Zinc.* B. B. zinc ores on charcoal form a thick coating of zinc oxide, yellow while hot, white on cooling. This, if treated with the cobalt solution and heated again, changes after cooling to a bright yellowish green.

Zr—*Zirconium.* Minerals containing zirconia, dis-

solved in HCl and the solution diluted with water, impart an orange-yellow color to turmeric paper.

THE USE OF THE KEY.

For economy of space the properties by which minerals may be determined are usually printed in a key, or in tables. To illustrate this, a short key to the twenty-four minerals whose descriptions have been given on pp. 202 to 225, is added. Take the specimen of *beauxite* for example.

We will suppose that you have found it, and do not know its name. Turn to the key. You notice that here all the specimens are divided into two classes. I. Those having metallic lustre. II. Those whose lustre is non-metallic. Your specimen falls in the second class, to which you will therefore turn at once. In this class you observe four divisions, A, B, C, and D, based upon the color of the streak. As your specimen has a pink streak and a reddish powder it must be referred to division B.

In this division you find three minerals distinguished by their action under the blow-pipe; and as you cannot melt even the smallest splinter of your specimen, although you use the hottest part of the flame, even the magnifying-glass revealing no trace of fusion on the thinnest edges or sharpest corners, you perceive that it must be either zincite or beauxite.

You now use the closed tube test. Heating some fragments in the c. t., you observe drops of water forming in the upper part of the tube. Therefore, the specimen must be beauxite.

To be doubly sure, however, you will turn to the reference list, page 253, and try some of the reactions for aluminum and iron.

Finally you will turn to the description of minerals —Number 13—for further evidence and confirmation. Determine all the specimens named in the key. Even if you already know their names, it will give you excellent practice, and prepare you to use the larger tables given in the standard books of reference.

One point remains to be noted: You will not always get in practice *pure* specimens; but there will often be a mixture of several minerals in one specimen unless you have the crystal. You must learn at once to make allowance for the presence of impurities in making your estimates of color, hardness, and all other physical properties. In case of a mixed specimen, it makes little difference which mineral you select for analysis. Give the preference, as a rule, to the one most largely and plainly developed; also in general select first metallic minerals. Study afterward to the best of your ability all other minerals that may appear in the specimen, always remembering that the chief benefit to you will come from the acquisition of the power and habit of careful and accurate observation, rather than from the mere knowledge of names which may result from that observation.

KEY FOR DETERMINING TWENTY-FOUR OF THE MORE COMMON MINERALS.

I. LUSTRE METALLIC.

(Some also Sub-metallic).

A.—Color Red.

		Name.	No. in Description.
1.	Malleable; reaction for Cu.....	Copper.	1

B.—Color Yellow.

(Sometimes tarnished on exposed surfaces)

1.	Magnetic before fusion; reaction for Fe,	Pyrrhotite.	4
2.	Magnetic after fusion.		
	a. Reaction for Fe and Cu.....	Chalcopyrite.	5
	b. Reaction for Ni.....	Millerite.	3

C.—Color Gray or Black.

1.	Magnetic before fusion; reaction for Fe with borax bead.		
	a. B. B. with soda on charcoal, white sublimate: reaction for Zn	Franklinite.	9
	b. B. B. with soda on charcoal, no sublimate; reaction for Ti	Menaccanite.	8
2.	B. B. magnetic after heating; Cr reaction;	Chromite.	10
3.	B. B. not magnetic: infusible; Mn reaction,.....	Pyrolusite.	12

II. LUSTRE NON-METALLIC.

A.—Streak or Powder Gray to Brown.

1.	B. B. fuses very easily, Ag reaction,.....	Cerargyrite.	6
2.	B. B. infusible; with soda on charcoal yields globules of Sn,.....	Cassiterite.	11

B.—Streak or Powder Red, Orange to Pink.

1. B. B. volatile; in c. t. black sublimate and Hg,.....Cinnabar. 2
2. B. B. infusible; in c.t. black while hot; on charcoal Zn reaction.....Zincite. 7
3. B. B. infusible; water in c.t.; Al and Fe reactions, ...Beauxite. 13

C.—Powder Green or Blue.

Reaction for Cu { *a.* Green,.....Malachite. 23
b. Blue, ...Azurite. 24

D.—Streak White or Very Light.

B. B. fuses easily, and
- *a.* Swells; becomes white; opaque; colors flame purplish,.....Spodumene. 14
- *b.* Swells to a whitish glass full of bubbles; alkaline reaction,.....Scapolite 18
- *c.* Swells to a greenish black globule,..Vesuvianite. 17
- *d.* Fuses without swelling to a brownish or greenish glass,.....Garnet, 16

2. B. B. white sprouts on surface; crimson flame,.....Strontianite. 22
3. B. B. small splinters fuse
   - *a.* Quietly to a white glass; not alkaline; gelatinizes with HCl,.....Elæolite. 19
   - *b.* With intumescence to a blackish mass; Ti reaction,.....Sphene. 20
4. B. B. almost infusible; gelatinizes with HCl; Zn reaction,
   - *a.* Yields water in c.t.....Calamine. 21
   - *b.* Yields no water in c.t.....Willemite. 15
5. infusible; does not gelatinize; yields water in c.t.; Al reaction,.....Beauxite. 13

If you are now desirous, as I trust you are, of con-

tinuing the study of minerals, you will need a book comprehensive enough to enable you to determine by the methods you have learned all known minerals of any importance. It gives me pleasure in closing to recommend Professor W. O. Crosby's Tables, for the Determination of Common Minerals, published by the author, Boston, ($1.00); and the Manual of Determinative Mineralogy, by Prof. George G. Brush, *John Wiley & Sons, New York*, [$3.50.]

LIST OF SPECIMENS PREPARED TO ACCOMPANY THIS BOOK.

1. Talc.
2. Gypsum.
3. Calcite, (*crystal of calc spar*).
4. Fluorite.
5. Apatite.
6. Orthoclase.
7. Quartz.
8. Cryolite.
9. Corundum.
10. Limonite.
11. Hematite.
12. Magnetite.
13. Cuprite.
14. Cast Zinc.
15. Carbon.
16. Sulphur.
17. Stibnite.
18. Prochlorite.
19. Quartz crystal.
20. White marble.
21. Sphalerite.
22. Celestite.
23. Ordinary stratified slate.
24. Roofing-slate.
25. Kaolinite.
26. Muscovite.
27. Biotite.
28. Halite.
29. Iron pyrites.
30. Galena.
31. Magnesium wire.
32. Graphite.
33. Serpentine.
34. Hornblende. (*Amphibole*).
35. Labradorite,
36. Chrysotile.

LIST OF PRINCIPAL ELEMENTS.

The following list comprises the elements found in

the minerals referred to in this book. The principal acid elements are printed in *italics*; some elements as arsenic, antimony, and manganese, form both acids and bases. *Oxygen* and *Hydrogen* unite both with basic-elements and acid-elements.

Name.	Symbol.	Atomic Weights.
Aluminum	Al	27 3
Antimony (Stibium)	Sb	122
Arsenic	As	75
Barium	Ba	137
Beryllium (Glucinum	Be (G)	9
Bismuth	Bi	208
Boron	B	11
Bromine	Br	80
Calcium	Ca	40
Carbon	C	12
Chlorine	Cl	35·5
Mercury (Hydrargyrum)	Hg	200
Molybdenum	Mo	96
Nickel	Ni	59
Nitrogen	N	14
OXYGEN	O	16
Phosphorus	P	31
Platinum	Pt	197
Potassium (Kalium)	K	39
Selenium	Se	79
Silver (Argentum)	Ag	108

Name.	Symbol.	Atomic Weights.
Chromium	Cr	52
Cobalt	Co	59
Copper (Cuprum)	Cu	63.4
Fluorine	F	19
Gold (Aurum)	Au	196
HYDROGEN	H	1
Iron (Ferrum)	Fe	56
Lead (Plumbum)	Pb	207
Lithium	Li	7
Magnesium	Mg	24
Manganese	Mn	55
Silicon	Si	28
Sodium (Natrium)	Na	23
Strontium	Sr	88
Sulphur	S	32
Tin (Stannum)	Sn	118
Tungsten (Wolfram)	W	184
Vanadium	V	51.4
Zinc	Zn	65
Zirconium	Zr	90

The *symbols* are abbreviations of the Latin names of

the elements; the *atomic weights* show the proportions, by weight, in which the elements combine. For instance: Galenite is lead sulphide PbS:

The atomic weight of Pb is . .	207
" " " S is . .	32
Together, ...	239

Thus, in every 239 grains, ounces, or pounds of galenite there are 207 grains, ounces, or pounds of lead, and 32 grains, ounces, or pounds of sulphur.

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